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OF THE
AMERICAN CHEMICAL SOCIETY.

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FEBRUARY, 1895.

EDWARD HART, Editor, J. H. LONG, EDGAR F. SMITH.
COMMITTEE ON PAPERS AND PUBLICATIONS.

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PAPERS ACCEPTED FOR PUBLICATION:

The Synthetic Food of the Future (President's Address). By H. W. Wiley.
On the Estimation of Sulphur in Pyrites. By G. Lunge.
Some Practical Points in the Manufacture of Nitroglycerol. By J. E. Blomen, Ph.D.
Coloring Matter in the California Red Wines. By W. D. Bigelow.
The Estimation of Tellurium in Copper Bullion. By Cabell Whitehead.
Methods for the Examination of Glycerol for use in the Manufacture of Nitroglycerol. By G. E. Barton.
Improved Methods of Water Analysis. By Irving A. Bachman.
Report of the Committee on Atomic Weights. By F. W. Clarke.
Note on the Estimation of Iron and Alumina in Phosphates. By K. P. McElroy.
The Superiority of Barium Hydroxide Solution as an Absorbent in Carbon Determinations in Steel. By J. O. Handy.
Condensation; and especially on the Condensation of Nitric Acid. By Edward Hart.
Some Points in the Distillation of Nitric Acid. By. Edward Hart.
Chemical Brick for Glover Towers. By Irving A. Bachman.
An Improvement on Danglers Laboratory Lamp (Note). By Andrew Lundteigen.
The Use of Sulphurous Acid ($HNaSO_3$) in the Manufacture of Glucose Syrup and Grape Sugar. By Horace E. Horton.
Note on the Purification of Glucinum Salts. By Edward Hart.
The Determination of Phosphoric Acid. By H. Pemberton, Jr.
Improvements in the Manufacture of Acetone. By E. R. Squibb.
On the Determination of Cane Sugar in the Presence of Commercial Glucose. By H. A. Weber and William McPherson.
On the Action of Acetic and Hydrochloric Acids on Sucrose. By H. A. Weber and William McPherson.

Issued with February Number, 1895.

Proceedings.

ANNUAL MEETING OF THE AMERICAN CHEMICAL SOCIETY.

The tenth general meeting of the American Chemical Society was held in Boston and Cambridge, Mass., Dec. 27th and 28th, 1894.

The Society was called to order at 10.20 A. M., Thursday, Dec. 27th, at the Massachusetts Institute of Technology, in room No. 22, of the Walker building.

Dr. T. M. Drown made a brief address of welcome to which President Wiley responded on behalf of the Society. Dr. Wiley expressed the special pleasure which the visiting chemists felt in meeting in the Massachusetts Institute of Technology, an Institution which has done so much for the advancement of science.

Mr. C. Tennant Lee, chairman of the Local Committee of Arrangements, made some announcements, and stated that on account of the inclement weather it would hardly be expedient to carry out the program which had been arranged for visits to the Experiment Station of the State Board of Health, and the Rubber Shoe Manufactory at Malden during the afternoon.

The minutes of the last general meeting were read and approved, and the minutes of the former meetings were also approved, as recorded by the General Secretary.

The General Secretary then read his annual report, which was as follows:

TO THE MEMBERS OF THE AMERICAN CHEMICAL SOCIETY:

GENTLEMEN: In presenting my annual report at the close of the present administration, a brief retrospective view of the organization and development of the Society, and a glance at its present condition will enable the members to gain instruction and encouragement from the past, hope and inspiration from the present, and a bright and confident outlook upon the future.

The American Chemical Society was the outgrowth of a

meeting of American chemists held in Northumberland, Pa., in August, 1874, to celebrate the centennial of the discovery of oxygen by Priestley. The Society was organized in 1876 and incorporated October 25, 1877. Its inception was vigorous, and those who took part in its organization had high hopes for the position it would occupy among the scientific societies of the world, for the inspiration and aid it would afford those who were engaged in chemical pursuits or in preparing others for such vocations, and for the ever increasing stimulus it would furnish to chemical science and industry throughout the American continent.

From the beginning the Society published regularly a Journal of its proceedings, including papers and discussions; many of the leading chemists of the country were enrolled in its membership; quite a number of prominent chemical manufacturers were among its associates; and the names of foreign chemists of distinction were upon its roll, either as active or as honorary members.

The following is a list of persons who have served the Society as its President:

*1876, J. W. Draper.	1886, A. B. Prescott.
*1877, J. Lawrence Smith.	1887, C. A. Goessmann.
1878, S. W. Johnson.	*1888, T. Sterry Hunt.
*1879, T. Sterry Hunt.	1889, C. F. Chandler.
*1880, F. A. Genth.	1890, H. B. Nason.
1881, C. F. Chandler.	1891, G. F. Barker.
1882, J. W. Mallet.	1892, G. C. Caldwell.
*1883, J. C. Booth.	1893, H. W. Wiley.
*1884, J. C. Booth.	1894, H. W. Wiley.
*1885, J. C. Booth.	

It is not until within the past few years, however, that the Society has fully justified the hopes of its founders that it would be in all respects a representative national organization. Its membership extended throughout the entire country and even included persons of other countries upon our own continent and in Europe, but, inasmuch as the meetings were all held in New York City, the members in other localities did not come into so active relation with the Society and its work as was desirable, and could not be prevailed upon to take sufficient interest in its

* Deceased.

welfare, nor to assume their share of responsibility in its management. Under these conditions there gradually developed among the members in other parts of the country the erroneous impression that the Society was local rather than national in its efforts and aims, and this impression was deepened by the fact that a number of prominent chemists throughout the country had either never been members of the Society or had withdrawn from active connection with it.

This state of affairs was unsatisfactory to every one, and it devolved heavy burdens upon the members residing in New York City and vicinity. There was no alternative, however; those members were obliged to carry the burden with heroic self-sacrifice until some plan could be devised by which the interest and active co-operation of a large body of chemists throughout the country could be permanently enlisted. To this end they labored with persistent determination and with unshaken faith in the ultimate result. The American Chemical Society had been recognized abroad for years as the organized body of the chemists of North America; it had already acquired something of prestige and established rights; and those who were made responsible for its management believed that it could be made to command the active support of the chemists of this continent, and to secure for them and their work adequate recognition among the chemists of the world.

The conferences, struggles, concessions, appeals to American chemists everywhere, whether members of the Society or not, and the hearty response to those appeals which was received from all over the country, need not be recounted here; but they will ever remain a noble example of self-sacrifice and true devotion to the welfare and progress of chemical science upon this continent.

The plan finally adopted by the chemists who had part in these counsels and efforts was to utilize what the American Chemical Society had already gained, to broaden and strengthen its foundation, and to effect radical changes in its management and operations. These changes were so outlined and provided for by the Society as to induce large numbers of prominent chemists, who had hitherto held aloof, to become members of the

Society and to add their influence, their energies, and their counsel to the forces that were already in operation to secure a complete and satisfactory reorganization of the Society.

Among the most effective measures adopted and continued up to the present time are the following:

(1) The management of all the important affairs of the Society is entrusted to a council of fifteen, composed of twelve members regularly chosen by ballot at the annual election of the Society—four elected each year, to serve for a period of three years—and in addition, the President, the General Secretary, and the Editor, ex-officio. The members of the council for the year 1894 are:

H. W. Wiley, Washington, D. C., President.
Albert C. Hale, Brooklyn, N. Y., General Secretary.
Edward Hart, Easton, Pa., Editor.
J. H. Appleton, Providence, R. I.
G. F. Barker, Philadelphia, Pa.
G. C. Caldwell, Ithaca, N. Y.
F. W. Clarke, Washington, D. C.
C. B. Dudley, Altoona, Pa.
W. L. Dudley, Nashville, Tenn.
J. W. Mallet, Charlottesville, Va.
Wm. McMurtrie, New York, N. Y.
T. H. Norton, Cincinnati, O.
A. B. Prescott, Ann Arbor, Mich.
Alfred Springer, Cincinnati, O.

(2) Local sections are established in different parts of the country, after the plan of the British Society of Chemical Industry. The presiding officers of the sections constitute the vice-presidents of the Society. Six local sections exist now, and there is a good prospect for the early formation of two or three more. Those already chartered are:

Rhode Island Section: Presiding Officer, Charles A. Catlin, 133 Hope St., Providence, R. I.; Secretary, E. E. Calder, Board of Trade Building, Providence, R. I.

Cincinnati Section: Presiding Officer, J. U. Lloyd, Court and Plum Sts., Cincinnati, O.; Secretary, E. C. Wallace, Room 71, Blymeyer Building, Cincinnati, O.

New York Section: Presiding Officer, Peter T. Austen, Polytechnic Institute, Brooklyn, N. Y.; Secretary, Morris Loeb, 37 East Thirty-eighth St., N. Y. City.

Washington Section: Presiding Officer, W. H. Seaman, 1424 Eleventh St., Washington, D. C.; Secretary, A. C. Peale, 605 Twelfth St., Washington, D. C.

Lehigh Valley Section: Presiding Officer, W. H. Chandler, South Bethlehem, Pa.; Secretary, Albert H. Welles, Lafayette College, Easton, Pa.

New Orleans Section: Presiding Officer, A. L. Metz, Tulane Medical College, New Orleans, La.; Secretary, Hubert Edson, Bartels, La.

(3) Two general meetings of the Society are held each year, the time and place of these meetings being determined by the council. General meetings have been held as follows:

1. Newport, R. I., August 6 and 7, 1890.
2. Philadelphia, Pa., December 30 and 31, 1890.
3. Washington, D. C., August 17 and 18, 1891.
4. New York, N. Y., December 29 and 30, 1891.
5. Rochester, N. Y., August 16 and 17, 1892.
6. Pittsburg, Pa., December 28 and 29, 1892.
7. Chicago, Ill., August 21 and 22, 1893.
8. Baltimore, Md., December 27 and 28, 1893.
9. Brooklyn, N. Y., August 15 and 16, 1894.

The change in membership during the past few years, since it was decided to arrange for general meetings of the Society to be held semi-annually in different localities, may be seen from the following statistics taken from the list of members published each year:

Nov. 1889,	169	members,	28	associates,	8	honorary	members.
Nov. 1890,	203	"	27	"	8	"	"
Nov. 1891,	269	"	24	"	9	"	"
Nov. 1892,	319	"	25	"	7	"	"
Sept. 1893,	423	"	30	"	7	"	"
Sept. 1894,	664	"	50	"	8	"	"
Dec. 1, 1894,	700	"	55	"	8	"	"
" 26, 1894,	720	"	55	"	8	"	"

Total, 783.

The present status as to membership should be given as follows:

Official roll of members	783
Elected but not required to qualify before Jan. 1, 1895	37
Applications now pending, no doubt of election	30
Total	850

By inspection of the foregoing figures it will be seen that the membership of the Society has more than doubled during the past two years, and has more than quadrupled since November, 1889.

(4) The Journal of the Society appears on the first of each month throughout the year, each number containing at least seventy-two pages, making a yearly volume of the minimum size of 864 pages. It has been found necessary, however, to enlarge the size of some of the numbers, and that for December, 1894, includes about 100 pages of papers and discussions in addition to proceedings, etc. It is the intention of the Society to make this the leading and representative chemical journal of the United States, and one of the most important scientific periodicals to be found anywhere. It will contain the papers read before the various sections of the Society and before its general meetings, together with such abstracts relating to the progress of chemical science and industry as may be necessary. A prominent feature will be its effort to represent industrial and commercial chemistry. Each branch of chemical science will receive due attention and support, and those engaged in any kind of chemical research, or in the application of chemistry to the arts, will find in the JOURNAL a wide scope for publication. Through this medium the followers of each branch of chemistry will be brought into intimate contact with each other, and will know also what those in other branches of chemistry are doing, and what progress is making in all directions.

During the year 1894, by order of the council, the General Secretary has been charged with the duty of collecting all the membership dues. A large amount of arrears has been collected, and we believe that the percentage of arrears at present is lower than ever before. The financial report of the General Secretary is presented in a separate statement.

At the beginning of the year, Edward Hart was reappointed editor, and the regular standing committees for the year were selected.

A committee, consisting of F. W. Clarke, Chairman, Albert C. Hale, Secretary, and W. L. Dudley, was appointed by the President, on authority of the council, to receive and consider

suggestions for amending the constitution, the main object sought being to have only one governing body for the Society, instead of two—the Council and the Board of Directors—as at present; this body to be selected from the membership without regard to the locality of their residence.

By the action of the Board of Directors the General Secretary was instructed to place on file a certificate of incorporation of the Society, signed by all of the directors, and so worded that the majority of the Board of Directors shall not be required to be residents of the State of New York. Upon presenting such certificate at the office of the Secretary of State in Albany, N. Y., it was rejected on the ground that, under the laws now existing such an amendment can not be made. Upon this, the Board of Directors, with the recommendation of legal counsel, took the following action: "*Resolved*, That the Board of Directors hereby request the American Chemical Society, at its annual meeting, to authorize the taking of such steps as may be necessary to secure the passage of a bill by the Legislature of the State of New York, which shall so change the present laws of the State as to allow the Society to choose its directors without regard to their being residents of the State of New York, or of any other State or locality."

At the summer meeting of the Society, held in Brooklyn, N. Y., August 15 and 16, 1894, Dr. Wolcott Gibbs, of Newport, R. I., was unanimously elected an honorary member of the Society. This is the first time an American has been elected to honorary membership. Upon being informed by telegraph of this action, Dr. Gibbs sent a reply expressing his sincere thanks to the Society, and his high appreciation of the honor they had conferred upon him. An engrossed certificate of his election, signed by every member of the council, has been forwarded to Dr. Gibbs in accordance with the provisions of the constitution.

During the year, charters have been granted by the Board of Directors for the formation of two new local sections, on Feb. 15, for the Lehigh Valley Section, and August 15, for the New Orleans Section. The charters for both of these have been forwarded and the Sections have been organized.

The outlook has never before been so promising as it is to-day.

The Society is now founded upon a broad and strong basis, and its mode of organization allows a ready modification of its management whenever experience may indicate that changes are desirable. The plan of holding general meetings twice a year, in different localities, strengthens the Society, rapidly extends its influence, and brings the chemists of the country into personal acquaintance with one another and into a knowledge of the vast amount of successful work which is done by them upon this continent. The formation of local sections deepens the interest felt in the general Society, and stimulates scientific activity and enterprise in those localities where such sections are established. The growth of the Society in numbers and influence has been almost phenomenal within the past few years, and when we consider its present attainments, and the number and standing of those who are now upon its roll of membership we may confidently assert that the American Chemical Society to-day takes rank with the leading organizations of its kind in the civilized world.

Respectfully submitted,

ALBERT C. HALE,
General Secretary.

BROOKLYN, N. Y., December 26, 1894.

FINANCIAL REPORT OF GENERAL SECRETARY OF AMERICAN CHEMICAL SOCIETY.

DECEMBER 1, 1894.

Dues collected for the year 1892		$ 15.00	
" " " " 1893		193.00	
" " " " 1894		2,851.00	
" " " " 1895		105.00	
Interest received on bank deposit		5.73	$3,169.73
Commissions deducted, 10% on $3,162.50		316.25	
Remitted to Treasurer	$2,655.00		
Balance in bank, December 1, 1894	198.48	2,853.48	
Total receipts			$3,169.73
Commissions deducted, 10% on $3,162.50			316.25
Amount due Treasurer:			
Account rebate to Local Sections	$30.33		
Commission overdraft F. W. Traphagen	.15	30.48	

Commission due General Secretary from dues received by Treasurer on General Secretary's account, 10% on $20.00	2.00	
Balance due Treasurer from General Secretary		28.48
Net commission of General Secretary		$287.77

After this the Treasurer's report was read by the Secretary and adopted, it having been previously audited by the Finance Committee.

AMERICAN CHEMICAL SOCIETY.

TREASURER'S REPORT FOR 1894.

Receipts.

Balance on hand, December 22, 1893	$ 74.05
Dues collected by Treasurer	622.09
Net dues and interest received from Secretary	2,853.48
Net commissions returned by Secretary	28.48
Cash received for subscriptions to Journal	575.42
" " " back numbers	159.48
" " " advertisements in Journal	131.80
" " " cancelled insurance policy	31.64
" " " Furniture sold to University C. N. Y.	100.00
	$4,576.44

Disbursements.

Expense of Journal and editorial office	$2,890.27
" " Secretary's office	413.12
" " Treasurer's office	24.72
" " Librarian's office	153.50
" " Baltimore meeting (balance)	4.50
" " Brooklyn meeting	60.55
General expense	16.50
Insurance	84.00
Rebate paid to Treasurer of N. Y. Local Section	200.00
" " " " " Washington Section	101.67
" " " " " Rhode Island Section	28.33
" " " " " Cincinnati Section	63.33
" " " " " Lehigh Valley Section	30.00
Balance on hand Dec. 22, 1894, in Bank of Metropolis, $505.19	
Postage stamps .76	505.95
	$4,576.44

CHAS. F. MCKENNA, *Treasurer.*

No bills or claims, presented up to date and audited, remain unpaid. Accounts examined and found correct, E. and O. E.

A. P. HALLOCK,
DURAND WOODMAN,
FRANK T. KING,
Finance Committee.

New York, Dec. 22, 1894.

Professor Edward Hart, Editor of the Society's publications, gave an oral report showing what had been done in connection with the publication of the Journal of the Society, and explained how the scope of the Journal could be widened and that periodical maintained at the highest standard of excellence.

The report of the Librarian was then called for, but as the Librarian was absent and no report had been sent in, Professors A. H. Sabin, of Long Island City, and Wm. McMurtrie, of New York, presented informal statements of the condition of the library. Upon the request of the President, the report of the Librarian was ordered to be spread upon the minutes when it should be received by the Secretary.

The Society then proceeded to the order of general business, and, upon the motion of Professor A. C. Hale, adopted the following:

Resolved, That the General Secretary of the American Chemical Society is hereby authorized and requested, to take such steps as may be necessary to secure the passage of a bill by the Legislature of the State of New York, which shall so change the statutes of said State as to allow the Society to choose its Directors without regard to their being residents of the State of New York, or of any other state or locality.

The arranged program was then taken up, and a paper entitled "Note on the Estimation of Iron and Alumina in Phosphates," by K. P. McElroy, was read by the General Secretary. Dr. Thos. M. Drown then followed with a description of the Lawrence Experiment Station of the Massachusetts State Board of Health, and the nature of the work which was carried on there for the purification of the sewage and water. He distributed among the chemists present, copies of two tables showing methods and results of the experiments of the Station. Dr. Drown mentioned four methods of filtration in use, that had been tried at the Experiment Station; namely, rapid filtration through gravel, aided by a current of air, chemical precipitation, sedimentation, and straining through coke. The result obtained through the first system was particularly valuable; no other method had been found so capable of rapidly purifying a large amount of sewage. By means of this method some six or seven hundred thousand gallons were purified daily, and the effluent

rendered as wholesome as ordinary spring water. Dr. Drown then spoke of the results which had followed the purification of the drinking water and pronounced them marvelous. In Lowell and Lawrence the death-rate from typhoid fever had been three times greater than the normal. The inhabitants were mostly operators in mills and obtained their drinking water from the Merrimac river, which was contaminated by the sewage of the cities and towns along its course; since the purification of the water by rapid filtration through gravel, the death-rate and number of cases from typhoid fever had been diminished to an unprecedented degree. The money for these experiments had been appropriated by the State of Massachusetts, and the result justified the confidence placed in the State Board of Health and its experiments. After some questions and discussions upon Dr. Drown's remarks, a paper entitled, "The Coloring Matter in California Red Wines," was read by W. D. Bigelow.

Two papers were read by title; namely, "An Examination of the Atmosphere of Cleveland, Ohio," by C. F. Mabery, and "Rainfall and Typhoid Fever," by Wm. P. Mason, Troy, N. Y.

Mr. J. O. Handy then read a paper on "The Advantages in the Use of Barium Hydrate Solution as an Absorbent in Carbon Determination in Steel."

After this Professor C. E. Munroe read a paper by J. E. Blomen, on "Some Practical Points on the Manufacture of Nitro-Glycerine," and a paper by himself upon "Tellurium from Copper Residues," and exhibited an ingot of tellurium prepared in the laboratory of the Columbian University, which is believed to be the largest specimen of that metal ever obtained. The ingot was about six inches in length, one and one-half inches in width, and something more than one-fourth of an inch in thickness, showing beautiful metallic luster and crystalline formation; it weighed 16.7 ounces; at the quoted price of this element the value of the ingot would be something over $250.

Professor Munroe then presented a paper, by Cabell Whitehead, on "The Estimation of Tellurium in Copper Bullion, and Professor Edward Hart read a paper entitled "Some Points in the Distillation of Nitric Acid." After some announcements, the Session adjourned at 1 P. M.

Upon the adjournment of the Session the chemists went directly to the Union Station, at Causeway St., where lunch was served by the Local Committee.

The weather having become fairly pleasant, it was decided here to take the train for Malden, after lunch, and visit the manufactory of the Boston Rubber Shoe Company, at Fells Station. Arriving at this destination, the chemists were very cordially received by the Superintendent, Mr. Robson, who escorted the party to the various departments in the building, carefully explaining to them all details of the various processes. They were shown first the washroom, where crude rubber from Para, Central America, and Africa, was sheeted and cleaned by running through the heavy rollers with a current of water pouring down upon it; next they were taken to the drying lofts, where the sheeted rubber is left for a few days exposed to the air at a temperature of 80°–90° F.; from this place the rubber is taken to the mixing room, where it is again run through heavy rollers, each having a different rate of speed; a mixture of vulcanizing agents (sulphur and litharge), coloring matter (lamp-black), inert filling (whiting), and reclaimed rubber and coal tar is then thrown upon the sheeted rubber as it passes between the rollers, which by their different rates of speed produce an effect like the kneading of dough, so that the material is thoroughly intermixed with the rubber, forming a homogeneous mass. The stock thus prepared is soft and extremely adhesive, its surfaces uniting firmly and permanently by contact and pressure. For convenience in handling, its adhesiveness can be temporarily removed by dusting the surfaces with flour, talc, etc.

The sheets formed by passing this material between the rollers can be prevented from sticking together by interposing plain sheeting for wrapping cloth, and they are transferred to the cutting room on cloth frames. All of the cutting is done on wooden blocks or zinc surfaces; the different methods of cutting are: *a*. Hand-knife and metal template. *b*. Hand die and mallet. *c*. Machine die. *d*. Machine knife (soles only). Fabrics are cut several thicknesses at once; sheet rubber singly.

Engraved surfaces, such as uppers and soles for rubber shoes, are made by passing the sheet rubber between rollers having

corresponding figures. Throughout all the processes of manufacture, except the making of heels, the rubber remains unvulcanized until everything else is done, vulcanization being the last step of all. Heels for heavy overshoes are made by hydraulic steam-jacketed presses, which form and vulcanize them at the same time.

In the boot and shoe rooms proper, the operations are as follows: *a*. Lasting innersole and lining over maple last. *b*. Applying the stay pieces, straps, etc. *c*. Covering outside with cloth or rubber upper. *d*. Attaching the soles. Each part requires to be rolled in place to insure adhesion and perfect contact. All dry surfaces are brushed over with rubber cement (solution of rubber compound in naphtha). Such goods as are to have a bright finish receive a coat of thin linseed-oil varnish previous to vulcanization.

The completed goods containing the lasts are vulcanized on iron cars, in chambers heated by steam coils. Temperature, 260°–270° F. Time, six to seven hours. The temperature is equalized and the volatile products removed by ventilation.

After vulcanization the lasts are removed and the boots and shoes packed for shipment.

SECOND DAY'S SESSION.

The meeting was called to order at 10.15 A. M. in the Chemical Lecture Room, Boylston Hall, Harvard University, Cambridge, Mass.

Professor C. Loring Jackson welcomed the chemists to Cambridge and the University; he described the location and arrangement of the various laboratories, museums, etc., which would be open for inspection after the morning session. President Wiley made some very appropriate remarks in response, after which the following papers were read:

Methods for the Examination of Glycerine for use in the Nitro-Glycerine Manufacture, by G. E. Barton. Read by title.

Report of the Committee on Atomic Weights for 1894, by F. W. Clarke. Read by the Secretary.

The Volumetric Determination of Phosphorus in Steel, by W. A. Noyes and J. S. Royse. Read by C. B. Dudley.

Note on the Manufacture of Zein, by A. H. Sabin.

The Volumetric Determination of Orcin, by H. S. Neiman. Read by title.

Note on the Purification of Glucinum Salts, by Edward Hart.

The Estimation of Sulphur in Pyrites, by George Lunge. Read by Edward Hart.

Some New Forms of Extractive Apparatus, by A. E. Knorr. Read by title.

Recent Work on the Atomic Weight of Oxygen, by E. W. Morley.

After Professor Morley's paper the Committee of Canvassers made their report, and the President announced the result of the annual election, which was as follows:

President—Edgar F. Smith, Philadelphia.
General Secretary—Albert C. Hale, Brooklyn.
Treasurer—C. F. McKenna, New York.
Librarian—F. E. Dodge, Brooklyn.
Directors (to serve two years)—H. W. Wiley, Washington; Wm. McMurtrie, New York; J. H. Appleton, Providence; A. A. Breneman, New York.
Councilors—(to serve three years)—F. W. Clarke, Washington; W. L. Dudley, Nashville; E. R. Squibb, Brooklyn; G. F. Barker, Philadelphia.

Upon the motion of Edward Hart, the Society unanimously passed a vote of thanks to the following individuals and corporations for courtesies received: Corporation and Faculty of the Massachusetts Institute of Technology; also Corporation of Harvard University; Dr. Thos. M. Drown; Professor C. Loring Jackson; Professor H. B. Hill; to the Boston Rubber Shoe Co., and Superintendent Robson; Authorities of the State Experiment Station at Lawrence, Mass., and members of the Associated Press. The address of the retiring President, "The Synthetic Food of the Future," was then presented by H. W. Wiley. At the close of the address, Dr. Wiley reviewed the history of the Society during the two years of his presidency, and expressed his thanks to the members, and especially to the various officers of the Society, who had aided him in making such marked advancement in numbers, in the character of the work and publications, and in the general reputation which the Society had gained throughout the world.

The Society then adjourned its session, and the chemists were tendered a lunch at the Colonial Club. The remainder of the

day was spent in examining the various laboratories, museums, and collections of the University.

The following named persons were registered in attendance at the meeting:

H. W. Wiley, Washington, D. C.
Edward Hart, Easton, Pa.
Albert C. Hale, Brooklyn, N. Y.
Charles L. Parsons, Durham, N. H.
Chas. W. Moulton, Poughkeepsie, N. Y.
C. H. White, Naval Laboratory, Brooklyn, N. Y.
T. M. Drown, Mass. Inst. Tech., Boston, Mass.
Chas. B. Dudley, Altoona, Pa.
Wm. McMurtrie, 106 Wall St., New York City.
Ellen H. Richards, Boston, Mass.
Arthur A. Noyes, Boston, Mass.
Augustus H. Gill, Mass. Inst. of Tech., Boston, Mass.
Walter S. Allen, Boston Gas Light Co., Boston, Mass.
William K. Robbins, Manchester, N. H.
Charlotte A. Bragg, Wellesley, Mass.
Mary E. Mathews, Painesville, Ohio.
Henry P. Talbot, Mass. Inst. Tech., Boston, Mass.
Walter L. Jennings, Boston, Mass.
Thomas Evans, Mass. Inst. Tech., Boston, Mass.
S. Lawrence Bigelow, Boston, Mass.
William T. Hall, 222 Marlborough St., Boston, Mass.
Rolfe M. Ellis, Hyde Park, Mass.
Chas. H. Miller, Baltimore, Md.
Chas. T. Pomeroy, Newark, N. J.
Durand Woodman, New York City.
Walter M. Saunders, Olneyville, R. I.
H. J. Wheeler, Kingston, R. I.
W. D. Bigelow, Washington, D. C.
Walter E. Smith, Providence, R. I.
Jas. O. Handy, Pittsburg, Pa.
Herbert L. Lamborn, Alliance, Ohio.
Henry H. Tozier, Haverhill, Mass.
Chas. W. Hapgood, Everett, Mass.
E. H. Laws, Bedford, Mass.
Sidney S. Emery, Philadelphia, Pa.
David B. Weston, Sharon, Mass.
Charles A. Meserve, Allston, Boston, Mass.
Albert Dunbar, Pittsfield, Mass.
C. Tennant Lee, Boston, Mass.
Bennett F. Davenport, Boston, Mass.

Edward W. Morley, Cleveland, Ohio.
Charles E. Munroe, Washington, D. C.
J. M. Crafts, Inst. of Technology, Boston, Mass.
J. H. Long, Chicago, Ill.
Severance Burrage, Inst. of Technology, Boston, Mass.
Frederick W. Howe, Durham, N. H.
Lewis H. Kittredge, Durham, N. H.
Charles A. Pitkin, South Braintree, Mass.
Charles A. Catlin, Providence, R. I.
George F. Andrews, Providence, R. I.
John H. Washburn, Kingston, R. I.
J. F. McGregory, Hamilton, N. Y.
J. W. Ellins, Framingham, Mass.
S. P. Sharples, Boston, Mass.
Walter Woodman, M.D., Cambridge, Mass.
A. H. Sabin, Long Island City, N. Y.
Albert M. Todd, Kalamazoo, Mich.
W. P. Cutter, Washington, D. C.
Howard C. Saunders, Providence, R. I.
Leonard P. Kinnicutt, Worcester, Mass.
Edw. P. Harris, Chester, Pa.
John B. Stearns, University of Vermont.
W. R. Orndorff, Ithaca, N. Y.
R. P. Williams, Boston, Mass.
Thomas B. Osborne, New Haven, Conn.
W. R. Livermore, Boston, Mass.
John F. Bannan, North Andover, Mass.
Horace A. Bishop, Providence, R. I.
Fred. L. Thayer, Providence, R. I.
Frank N. Smalley, Westboro, Mass.
S. D. Gage, Mass. Inst. Tech., Boston, Mass.
Harris E. Sawyer, Cambridge, Mass.
George Oenslager, Harrisburg, Pa.
Martin H. Ittner, St. Louis, Mo.
Frederick L. Dunlap, Chillicothe, Ohio.
Delia M. Stickney, Cambridge, Mass.
C. Belle Kenney, East Boston, Mass.
G. C. Hodges, Utica, N. Y.
A. E. Dolbear, Tufts College, Mass.
C. L. Jackson, Cambridge, Mass.
H. B. Hill, Cambridge, Mass.
Fred. W. Morse, Durham, N. H.
W. B. Lindsay, Dickinson College, Carlisle, Pa.
Stephen Freeman, Boston, Mass.

COUNCIL.

The bill for printing the December number of the JOURNAL, amounting to more than $250, it was submitted to the Council according to the Constitution, and approved by them January 5, 1895.

NEW MEMBERS ELECTED DECEMBER 26.

Becher, Jesse B., Lincoln, Neb.
Brown, Lucius P., Nashville, Tenn.
Chase, Geo. W., Phillipsburg, N. J.
Franklin, Edward Curtis, Lawrence, Kan.
Grindly, Dr. H. S., East Green St., Champaign, Ill.
Harrison, Herbert E., B.S., Agricultural College, Mich.
Hiltner, Robert S., Lincoln, Neb.
Meeker, G. H., Longdale, Va.
Miller, Hugh L., Wilmington, N. C.
Nelson, E. K., 6814 Union Ave., Englewood, Ill.
Nicholson, Edward E., Lincoln, Neb.
Rossman, William L., B.S., Agricultural College, Mich.
Wardman, Geo. B., care of Detroit Copper Company, Morenci, Ariz.
White, A. H., West University Ave., Champaign, Ill.
White, John, Ph.D., Lincoln, Neb.

ELECTED JANUARY 12.

Barrows, W. A., Jr., Sharpsville, Pa.
Beeson, J. L., care Audubon School of Sugar, Audubon Park, New Orleans, La.
Billings, Edgar F., 165 High St., Boston, Mass.
Fassett, Chas. M., Spokane, Wash.
Frohman, Edward D., 68 W. Third St., Cincinnati, Ohio.
Graham, Robert O., Bloomington, Ill.
Guess, George A., M.A., Fairview via. Penticton, B. C.
Homburg, Fred., 40 Clifton Ave., Cincinnati, Ohio.
Loewenstein, Harry, 144 Hackberry St., Cincinnati, Ohio.
Mather, Henry A., 27 Pierrepont St., Brooklyn, N. Y.
Neher, Fred., Princeton, N. J.
Peyton, William C., care The Cal. Powder Works, Santa Cruz, Cal.
Roth David M., 36 Ingleside Place, Cincinnati, Ohio.
Shaw, Professor G. W., Corvallis, Benton Co., Oregon.
Smith, J. D. Audley, Clifton, Ariz.
Swenson, M., Monadnock Block, Chicago, Ill.
Thorburn, A. D., 465 State St., Chicago, Ill.
Tucker, Francis F., Nebraska City, Neb.
Wrampelmeier, Theodore J., Custom House, San Francisco Cal.

ELECTED JANUARY 17.

Freeman, Stephen, 79 Montgomery St., Boston, Mass.
Guild, Frank N., Greenboro, Vt.
Johnson, Emil F., 96–98 Maiden Lane, N. Y. City.
Miller, Edmund H., Ph.D., 41 East 49th St., N. Y. City.
Newell, H. E., Saxtons River, Vt.
Pickert, Lee W., Amer. Sugar Ref. Co., Granite St., South Boston, Mass.
Redding, Allen C., Findlay, Ohio.
Spaulding Edward G., 43 W. 32nd St., N. Y. City.
Worcester, Charles P., State House, Boston, Mass.

ASSOCIATE ELECTED JANUARY 17.

Fosdick, Eugene H., San Diego, Cal.

CHANGES OF ADDRESS.

Barton, G. E., Hatfield, Mass.
Catlin, Chas. A., 133 Hope St., Providence, R. I.
Clarke, F. W., U. S. Geological Survey, Washington, D. C.
Dickerson, E. N., 253 Broadway, N. Y. City.
Dudley, Chas. B., Drawer 334, Altoona, Pa.
Eichberg, Julius H., M.D., 50 West 9th St., Cincinnati, Ohio.
Hahn, A. G. C., care of Nepera Chemical Company, Nepera Park, N. Y.
Jackson, C. E., care of Wm. S. Merrell Chemical Company, Cincinnati, Ohio.
Kedzie, Frank S., 414 Ottawa St., Lansing, Mich.
Lanning, John G., care of Cleveland R. M. Company, Cleveland, Ohio.
Latham, Woodville, Hotel Bartholdi, Madison Square, N. Y. City.
Randolph, R. B. F., Hoagland Laboratory, Brooklyn, N. Y.
Smith, Herbert E., Yale Medical School, New Haven, Conn.
Wardman, Geo. B., care of John L. Elliot, Ahualulco, Estado de Jalisco, Mexico.

ADDRESSES WANTED.

Davis, G. C., old address, 311 W. Mahantonga St., Pottsville, Pa.
Ray, Victor, old address, Anaconda Electrolytic Refinery, Anaconda, Montana.

MEETINGS OF THE SECTIONS.

WASHINGTON SECTION.

Thursday, December 13, 1894.

The Society was called to order at 8 P. M. by the President, Dr. Seaman, with twenty members present. The minutes of the 76th meeting were read and approved.

The first paper of the evening was read by Mr. Carr on "The Dehydration of Viscous Organic Liquids," by Oma Carr and F. T. Sanborn.

The authors reviewed briefly the various accepted methods for the dehydration of organic liquids, noting the peculiarly conflicting statements of authorities regarding oxidation and decomposition during heating. Statements of Jungfleisch and Lefranc, Dieck and Tollens, Battmann, and others are not in accord regarding the decomposition of levulose at various temperatures.

The authors discussed the essential factors of the ideal dehydration method, stating these to be: (1) Homogeneity of liquids. (2) Uniform distribution of small quantities over large surfaces. (3) Uniform temperature. (4) Rapidity of dehydration.

"It has been found that as these conditions have been more perfectly attained, results have been more generally concordant. It may be said that, given a homogeneous liquid and uniformity of temperature, the method which insures wider distribution will provide greater rapidity of drying. Inasmuch as at a given temperature, alteration of the material will be proportionate to the time, results closer to the truth will be afforded by the method accomplishing the drying in shorter period."

The method devised by the authors is in accord with the above principles, and has been applied, so far, to the following range of materials: Heavy sorghum molasses, mixtures of sugar and levulose, milks, cream, maize juice (22 years old), beet juices, honies, sorghum juices, commercial glucoses, and rosin turpentine solutions. The work on these materials covers over 4,000 determinations, in which the maximum variation in quadruplicates was 0.09 per cent., minimum 0.0, and mean 0.057 per cent. These results were obtained in from two to four hours, according to the character of the material.

The method is based upon the porosity of pumice stone, the

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Lanning, John G., care of Cleveland R. M. Company, Cleveland, Ohio.
Latham, Woodville, Hotel Bartholdi, Madison Square, N. Y. City.
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The method is based upon the porosity of pumice stone, the

latter being crushed into fragments the size of a pea, strongly ignited, cooled, and the liquid to be dried, weighed from a pipette-weighing bottle, and uniformly distributed over the stone. The drying is conducted as usual.

The authors reserve the details of the method for future papers, work of a more complex character being now under way.

Discussion.

Professor Wiley said the method had been proved accurate in giving duplicate results, although agreement is not proof of accuracy. The method could be used to determine matters volatile at 100°, chiefly water, as can be proved by passing them through sulphuric acid. Many of the substances to be dealt with contain levulose, and it is remarkable that we don't know just what takes place upon drying it, as the most opposite statements have been made. The determination of water is one of the most difficult things in agricultural chemistry, and this method is a step in advance. He referred to the fact that by the use of aluminum, flat dishes of wide expanse can be obtained.

Mr. Cutter stated that he had used practically the same method in the Utah Station, except that he had used the Hoffmeier capsules, and that the material, after being dried, could be utilized in other ways, especially the determination of fats in the analyses of meats. In his opinion, correct sampling of green material and the correct determination of moisture in original samples was, at present, one of the most perplexing problems with which agricultural chemists have to deal.

Professor Wiley said that green fodders should be shredded. The Laboratory of the Agricultural Department has the only shredding machine in the United States and it furnishes a good pulp.

Professor Seaman asked if the idea was to obtain a sufficient surface of pumice stone and was answered affirmatively.

W. D. Bigelow and E. E. Ewell called attention to a continuous extractor for large quantities of material. The apparatus consists of a percolator with slightly diverging sides, supported by an iron tripod, which in turn is placed in a shallow tray of enameled iron, or some other material which is not attacked by mercury. A crystallizing dish is placed below the percolator, and a bell-jar, with a reflux condenser attached, is placed over

the percolator, with its edges resting on the tray. The tray is then partially filled with mercury to prevent the escape of the ether.

January 10, 1895.

President W. H. Seaman in the chair; 14 members present; The following officers were elected:

President—Chas. E. Munroe.
Vice-Presidents—E. A. de Schweinitz and W. D. Bigelow.
Treasurer—W. P. Cutter.
Secretary—A. C. Peale.
Additional Members of Executive Committee.—H. W. Wiley, F. P. Dewey, F. W. Clarke, and W. H. Seaman.

CINCINNATI SECTION.

December Meeting.

Messrs. W. G. Wallace, Richard W. Proctor, George F. Feid, and Charles E. Jackson were elected to membership in the Section.

The election of officers for 1895 resulted as follows:

President—Karl Langenbeck.
Vice-Presidents—B. D. Westenfelder, and Irwin J. Smith.
Treasurer—Henry B. Foote.
Secretary—E. C. Wallace.
Directors—Dr. S. P. Kramer, Professor O. W. Martin, and H. L. Nickel.

Papers were read by Dr. S. P. Kramer on "Diphtheria Antitoxin," and Dr. Alfred Springer on "Elective Fermentation in Diabetes."

The following resolution was adopted:

Resolved, That the Cincinnati Section of the American Chemical Society respectfully recommends to the Committee on Nomenclature and Spelling of the JOURNAL of the Society, the use of the rules concerning the spelling of chemical terms recently adopted by the American Association for the Advancement of Science, limiting, however, the application of the rule concerning the spelling of *gramme* to articles in the field of medical or pharmaceutical chemistry.

Tuesday, January 15, 1895.

Vice-President Westenfelder in the chair.

Mr. F. Homburg and E. D. Frohman, recently elected members of the American Chemical Society, were elected to membership in the Section.

Mr. E. Twitchell read a paper on "The Separation of Solid and Liquid Fatty Acids." After reviewing the different methods of analysis of fats, Mr. Twitchell stated, that in his opinion, by precipitating all the solid and part of the liquid fatty acids from alcoholic solution by lead acetate (Jean's method) and determining iodine in a number of both original fatty acids and liquid acids left in the alcoholic solution, by calculation therefrom the amount of saturated and unsaturated fatty acids present in the fat could be determined with a fair degree of accuracy.

Dr. H. E. Newman read a "Report of Progress in Organic Chemistry," touching briefly upon L. Gatterman's method of reducing aromatic nitro compounds by a current of electricity; B. Tollens' work on the birotation of glucose; the synthesis of polyhydric alcohols by B. Tollens, H. Hosaus, and others, and Trillat's method of manufacturing dyes from formaldehyde and several amines, etc.

Reports of the Treasurer and Secretary, for 1894, were read and adopted. The Secretary's report showed an increase of six in membership, and that twenty papers had been read at the eight meetings held during the year.

LEHIGH VALLEY SECTION.

Thursday Evening, January 17, 1895.

The annual meeting was held at the Franklin House, Easton, Pa., six being present. A pleasant social time was spent, the only business transacted being the election of officers for the coming year. Messrs. Green and Brinker acted as tellers, and reported, on opening the ballots, the election of the following:

Presiding Officer—Edward Hart.
Secretary—Albert H. Welles.
Treasurer—Albert L. Colby.
Executive Committee—George P. Scholl, and J. W. Richards.

VOL. XVII. FEBRUARY, 1895. NO. 2.

THE JOURNAL

OF THE

AMERICAN CHEMICAL SOCIETY.

A MODIFIED ARRANGEMENT OF THE ELEMENTS UNDER THE NATURAL LAW.

BY F. P. VENABLE.

Received December 12, 1894.

A quarter of a century has passed since the first announcement of the Natural Law and the publication of Mendelejeff's table. The truth of the law, in a general way, seemed to be accepted very readily by chemists. It was incorporated in text-books and there explained, but comparatively little use has been made of it in teaching the science. Even Mendelejeff himself, in his Principles of Chemistry, has not made the fullest use of it. Victor Meyer, in his lecture before the German Chemical Society more than a year ago, showed how it might be used, and how he used it himself, and, probably, this will do much towards popularizing its use.

There must be some reason why so great a help to scientific study, is not made more use of. Does it lie in a lingering distrust of the law itself or failure to accept it, or is it because of the imperfections in the arrangements of the elements offered by Mendelejeff and others? It is most probably due to the latter, and this paper is presented with the hope of clearing up some of these difficulties.

The modern chemical world has recognized in the discovery of Mendelejeff, the greatest step forward since the announcement of the Atomic Theory. It is too much to expect that so great a discovery should spring full-panoplied from the head of its author.

Doubtless many have observed the imperfections of the law's original form, or rather the table as first given out. Probably some have ventured to comment upon it. Such criticisms have, however, escaped me with one or two exceptions. It is with much hesitation, that I venture to point out what seem to me imperfections and blemishes, in so great a work. Few may agree with me in calling them imperfections. I do not purpose to detract one particle from the greatness and importance of the essential truths contained in this discovery. Mendelejeff's table, as we have it at present, is a great advance upon the first one published by him in 1869, which must be pronounced tentative only, and decidedly unsatisfactory. The table of Victor Meyer is far behind it in presenting the facts of the periodic law. There have been many attempts at devising a graphic representation of this law. I know of none which can be called a real aid to the student, or which do not introduce new ideas which, to say the least, have no basis in the facts as known to us at present. None of them can be regarded as a safe substitute for the simple table of Mendelejeff.

Taking this table I would venture to point out some obstacles to its full acceptance. These have been in part revealed to me by the effort at a presentation of these truths of nature to honest-minded, clear-sighted young men.

Before mentioning the difficulties which lie here in the path of a teacher, I must say, by way of preface, that my criticisms, are aimed at what I may be allowed to call the unessentials of the law. Mendelejeff's great feat was in seeing clearly, and announcing intelligently, that the properties of the elements are dependent upon and determined by the atomic weights. This is the essential of the Natural Law and is in accord with our fullest knowledge. The second part of the law, as usually stated, that these properties are periodic functions, attempts, in a measure, to define the dependence. It may also be true, but it is not fully proved and it is open to objections. It seems to me that this hypothetical portion could well be left in abeyance until fuller knowledge gave it a stronger footing, meanwhile substituting something less open to criticism, and which can not weaken the central truth.

TABLE I.

Group.	I	II	III	IV	V	VI	VII	VIII
Series 1	H	—	—	RH_4	RH_3	RH_2	RH	{ Hydrogen compounds.
" 2	Li .	Be .	B .	C .	N .	O .	F .	
" 3	. Na	. Mg	. Al	. Si	. P	. S	. Cl	
" 4	K .	Ca .	Sc .	Ti .	V .	Cr .	Mn .	Fe . Co . Ni . Cu
" 5	. (Cu)	. Zn	. Ga	. Ge	. As	. Se	. Br	
" 6	Rb .	Sr .	Y .	Zr .	Nb .	Mo .	— .	Ru . Rh . Pd . Ag
" 7	. (Ag)	. Cd	. In	. Sn	. Sb	. Te	. I	
" 8	Cs .	Ba .	La .	Ce .	Di? .	— .	— .	— — — —
" 9	. —	. —	. —	. —	. —	. —	. —	
" 10	— .	— .	Yb .	— .	Ta .	W .	— .	Os . Ir . Pt . Au
" 11	. (Au)	. Hg	. Te	. Pb	. Bi	. —	. —	
" 12	— .	— .	— .	Th .	— .	U .	— .	
	$\frac{R_2O}{——}$	$\frac{R_2O_2}{RO}$	$\frac{R_2O_3}{——}$	$\frac{R_2O_4}{RO_2}$	$\frac{R_2O_5}{——}$	$\frac{R_2O_6}{RO_3}$	$\frac{R_2O_7}{——}$	Higher Oxides. RO_4

TABLE II.

Groups.	Higher salt-forming oxides.	Typical, or first small period.		Large periods.				
				1st.	2nd.	3rd.	4th.	5th.
I	R_2O		Li 7	K 39	Rb 85	Cs 133	—	—
II	RO		Be 9	Ca 40	Sr 87	Ba 137	—	—
III	R_2O_3		B 11	Sc 44	Y 89	La 138	Yb 173	—
IV	RO_2		C 12	Ti 48	Zr 90	Ce 140	—	Th 232
V	R_2O_5		N 14	V 51	Nb 94	—	Ta 182	—
VI	RO_3		O 16	Cr 52	Mo 96	—	W 184	U 240
VII	R_2O_7		F 19	Mn 55	—	—	—	—
VIII	}			Fe 56	Ru 103	—	Os 191	—
	}			Co 58.5	Rh 104	—	Ir 193	—
	}			Ni 59	Pd 106	—	Pt 196	—
I	R_2O	H = 1	Na 23	Cu 63	Ag 108	—	Au 198	—
II	RO		Mg 24	Zn 65	Cd 112	—	Hg 200	—
III	R_2O_3		Al 27	Ga 70	In 113	—	Te 204	—
IV	RO_2		Si 28	Ge 72	Sn 118	—	Pb 206	—
V	R_2O_5		P 31	As 75	Sb 120	—	Bi 208	—
VI	RO_3		S 32	Se 79	Te 125	—	—	—
VII	R_2O_7		Cl 35.5	Br 80	I 127	—	—	—
			2nd small period.	1st.	2nd.	3rd.	4th.	5th.

Take the tables from the first volume of Mendelejeff's Principles of Chemistry and examine them. First, we find two kinds of periods made use of—periods containing seven elements, and those containing seventeen. These latter are divided into sevens and threes. If it had only been possible to arrange all of the elements in sevens, as Newlands attempted to do, the periodic idea would have been most convincing, and the Law of Octaves, running through nature, would have seemed most wonderful. But these elements do not admit of being arranged in this way, and the use of periods of different lengths, is to a fresh young mind, unacquainted with mathematical expedients, somewhat forced.

Secondly, there is a very anomalous position assigned to the triads, or as sometimes written, the tetrads, Fe, Co, Ni, Cu, etc. They have been set off to themselves, clearly so as to make the other elements fall, even approximately, into their places, and into the proper sevens. I say approximately, for the student soon sees that although there is a similarity, there is also a wide difference between the elements of the first seven and the last in any period of seventeen.

Thirdly, in the lower periods, in order to get elements to fall into their places, a great many unknown elements have to be interpolated. Thus between cerium and ytterbium, the next element in the list, there are blank places for sixteen elements. The third large period of seventeen has only four known elements in it, and the fifth has only two. That means that here we have a period actually constructed out of fifteen unknown elements and two known ones. This exceeds some of the triumphs of geology in the construction of skeletons of extinct animals. Of the five periods, only one is completely filled out. To say the least, this shows a very imperfect knowledge of the elements, or a great deal of guess work. In the table there are sixty-four known elements and thirty-five blanks for elements yet to be discovered. I hardly think it possible that the majority of chemists, believe that, after all of the diligent search for the past century, less than two-thirds of the elements have been discovered. Where are the others in hiding? Will they be discovered by the spectroscope among the rare earths? There is certainly hope of finding some of them there, but the number which this

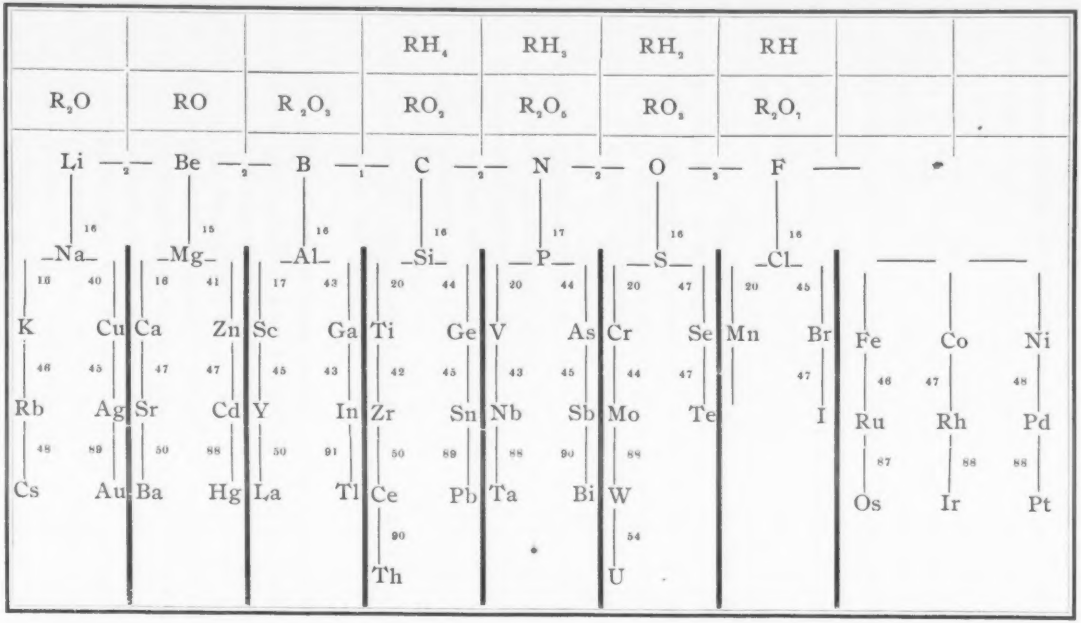
statement of the law would require us to find, is simply appalling. The average student thinks, in all honesty, that the coincidences of the first part of the table, will scarcely justify such forcing and wholesale interpolation. If our knowledge of the elements be so imperfect as that, we have no right to force them into periods, in fact, any law based upon their atomic weights would be of the most tentative character, and likely at any time to be overthrown by the discovery of the lost or unknown ones. How do we know that the big one-third, now unknown, may not upset all calculations, when found? Such a law ought rather to be called a working hypothesis. We are venturing a great deal upon a very imperfect knowledge of the ones we have in hand. For if one reckons up the number of elements, for which we have satisfactory determinations of the ato mic weights, he will find that they are less than forty. The periodic idea may be true, but we do not know enough about these elements yet, to be able to give this idea a very prominent place in the Natural Law, and we ought to avoid the assumption of so many unknown elements unless it is absolutely necessary.

As I do not intend to tear down without some effort at rebuilding, I would, with much real diffidence, for I realize that I may be looked upon as one who would rush in where only the great masters of the science can safely tread, offer the following suggestions.

The first suggestion is that the wording of the Natural Law be so changed as to read: "The properties of the elements are dependent upon and determined by the atomic weights." The somewhat difficult idea of functions is simplified and periodicity is subordinated. Then the following table might be substituted for the one ordinarily given.

It is not greatly changed and not much originality is claimed for it, but, however slight the changes, I would insist upon their value, because they do away with the dependence upon periods and they certainly make the table easier, more intelligible, and more useful to the student. No very doubtful element is included in it. There is room for additional elements as discovered, but the table is not dependent upon them. Lastly the inter-relation is more clearly brought out.

TABLE III.


RH4 RH3 RH2 RH
R2O RO R2O3 RO2 R2O5 RO3 R2O7
Li —2— Be —2— B —1— C —2— N —2— O —2— F —
16 15 16 16 17 16 16
Na Mg Al Si P S Cl
16 40 16 41 17 43 20 44 20 44 20 47 20 45
K Cu Ca Zn Sc Ga Ti Ge V As Cr Se Mn Br Fe Co Ni
46 45 47 47 45 43 42 45 43 45 44 47 47 46 47 48
Rb Ag Sr Cd Y In Zr Sn Nb Sb Mo Te I Ru Rh Pd
48 89 50 88 50 91 50 89 88 90 88 87 88 88
Cs Au Ba Hg La Tl Ce Pb Ta Bi W Os Ir Pt
90 54
Th U

I do not maintain that this table could ever have been discovered without the idea of periods, though I can see no reason why it might not. The periods still underlie it, but they are out of sight for the present, and are not necessary. The table is not dependent upon them.

The table is constructed as follows. There are seven group elements, having a mean increment of two in their atomic weights. It is by no means essential that there should be just seven of these. At present we do not know more, but I think there is possibly a place for one more, having the atomic weight twenty-one and differing widely from the others as it occupies a singular position.

These group elements are also to be called bridge elements, as they show marked gradation of properties from one to another and serve to bridge over the groups and connect one with the other. Linked to them by an increment of sixteen, are seven typical elements. These show the distinctive properties of the groups to which they belong and a wider divergence from the next group to them. From them can be deduced the properties of the remaining elements of the group. Thus, in group I, Li is the bridge or group element, and Na the type. From this type two lines of elements diverge, averaging three to the line. These triads would, of course, be changed into tetrads or pentads by the discovery of more elements. No importance can be attached to the fact that at present they are triads. There is a distinct increment for each line of elements. These can be averaged thus:

Fig. 1.

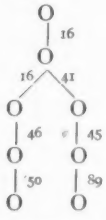


Fig. 2.

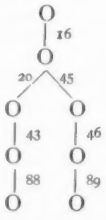


Figure 1 represents the arrangements and increments for the first three groups, and figure 2, that in the last four groups,

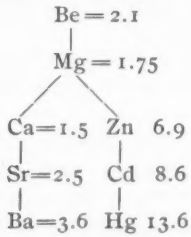
the increments showing a variation. These increments could be averaged in all except one case, and the agreements with known atomic weights would be close enough to admit of the easy arrangement of the elements in the prescribed order. Naming the triads right triad and left triad, respectively, we find that these averaged increments would be as follows: The increment from group to type element is sixteen; from type to first element in left triad is eighteen; to second element in left triad is sixty-three; to third element in left triad is 112. On the other side to first element in right triad is forty-four; to second element in right triad is eighty-nine; to third element in right triad is 177.

The one exception mentioned is in the increment from type to third element in right triad, in groups IV to VII. Instead of being 112, the increment here is 141.

To the right of group VII we have three triads which have the regular increments belonging to the left triads, namely, forty-seven and eighty-eight. They are without any type element. It seems most likely that they belong to one group. The group element would have an atomic weight of twenty-one, and the type one of thirty-seven.

The arrangement in the table then, is partly one based upon regular increments in the atomic weights, and, since these weights are but poorly known, partly upon our knowledge of the properties of the elements. When it is recalled that about one-half of the atomic weights are imperfectly known, it will be evident that these averaged increments are approximations only. It is impossible to bring out such perfect symmetry as is obtained in the homologous series in organic chemistry. And yet these groups should be something of the same kind. Following the analogy to the organic hydrocarbons a little further, may not the existence of an element in two different conditions as to valence, etc., as, for instance, copper, or mercury, or iron, be looked upon as a species of isomerism? Such speculations are of little use, however, and quite apart from our present purpose. I have found this table very useful in teaching elementary chemistry, and it can most profitably be made the basis of the entire course. Thus, in the first four groups, the left triads contain

the elements most closely resembling the types. In the last three they are to be found in the right triads. As to natural occurrence of the elements, in the first four groups those in the left triads occur in the same compounds, and generally in connection with the type; those in the right triads occur as the type, or as sulphides, or free. In the last three groups this is reversed. The right triad elements occur as the types, and the left triad elements as the type or as oxides. So, too, the properties of the elements show this relation to the types. Take as an example the specific gravities in group II.



It is not necessary to pursue this part of the matter at greater length. The careful teacher will easily work out all of these comparisons for himself, and will find that chemistry taught by the table is shorter (so much repetition being saved) and is easier for the pupil, and its symmetry and beauty are much more easily brought out. There is no special claim for originality made here. The germs of such a table, or arrangement, can be found in several text-books, but I do not know of any in which the idea is fully developed, or such a table as this is given.[1] I offer the whole as a suggestion. Perhaps some may find it useful who have met the same difficulties which I have encountered. Others may have overcome these difficulties in a still better way than this. I think, at least, all will agree that there are difficulties, and very serious ones, in the use of Mendelejeff's table, or that of Victor Meyer, as given by their respective authors.

[1] The arrangements of Bayley, Hinrichs, and Wendt are somewhat similar, but the ideas which I would make prominent, are obscured by other considerations and speculations.

THE DETERMINATION OF POTASH IN KAINITE.

By Rudolf de Roode.

Received December 12, 1894.

AN aqueous solution of kainite contains nothing except potash which, upon evaporation with platinum chloride, forms a compound which is insoluble in alcohol and the ammonium chloride wash ordinarily used in potash determinations. It is, therefore, quite superfluous to precipitate out anything from such a solution by such reagents as barium chloride, ammonium oxalate or carbonate, etc., and to evaporate and ignite the residue obtained. A most accurate result is obtained by evaporating an aliquot portion of a filtered aqueous solution of kainite directly with platinum chloride. Upon the addition of alcohol to the residue obtained by evaporation, nearly to dryness, with platinum chloride, the sodium sulphate present renders the potassium platinochloride sticky and difficult to wash, just as in the Lindo-Gladding method. The sodium sulphate is readily washed out with the ammonium chloride wash, however, leaving a pure potash double salt. The method of direct evaporation of a solution of kainite with platinum chloride has the advantage that there is no loss from occlusion or retaining of potash by precipitates, since no precipitants, such as barium chloride, ammonium oxalate, etc., are used. There can also be no loss from spattering, volatilization, etc., upon ignition, since no ignition is used. Using the same aqueous solution of a sample of kainite, and making determinations by the Lindo-Gladding, the Alternate, and the Stassfurt methods, and also by direct evaporation with platinum chloride, I obtained the following results:

Lindo-Gladding	12.36	per	cent.	potassium	oxide.
Alternate	12.42	"	"	"	"
Stassfurt	12.40	"	"	"	"
Direct evaporation with platinum chloride	12.54	"	"	"	"

Agricultural Experiment Station,
Morgantown, West Virginia,
December 10, 1894.

THE OXIDATION OF ORGANIC MATTER AND THE DECOMPOSITION OF AMMONIUM SALTS BY AQUA REGIA, IN LIEU OF IGNITION, IN THE DETERMINATION OF POTASH IN FERTILIZERS.

By Rudolf de Roode.

Received December 12, 1894.

IN the determination of potash in fertilizers errors may arise, by reason of a loss of potash, when the residue obtained upon evaporation is ignited to destroy organic matter and to drive off salts of ammonium. Such loss may easily occur, either by spattering or a mechanical carrying off of potassium sulphate by the voluminous fumes of sulphuric acid used in the Lindo-Gladding method. In the alternate method, a loss of potash may occur by reason of direct volatilization. Extreme care must be exercised in the ignition, both in the Lindo-Gladding, and in the alternate method, in order to obtain accurate results.

In order to avoid the errors which might arise from such ignition, and, in fact, to do away with the necessity for any ignition whatever, I tried the action of aqua regia. J. Lawrence Smith has shown that salts of ammonium are readily destroyed by aqua regia, and it is well known that organic matter is oxidized thereby. I found, upon evaporating a solution of a mixed fertilizer, prepared as in the Lindo-Gladding method, with nitric and hydrochloric acids, that the salts of ammonium were completely destroyed, and that the organic matter was either entirely destroyed, or else so highly oxidized that it no longer exercised a reducing action upon platinum chloride.

By preparing a suitable solution of a fertilizer and evaporating an aliquot portion of this solution with aqua regia, we ought, therefore, to obtain a residue quite as well adapted to the determination of potash by platinum chloride as a residue obtained by evaporation and ignition.

The *details* of a method for the determination of potash in fertilizers, based upon this principle, have not yet been thoroughly worked out. Some excellent results have been obtained and

some results which were not so good, but, on the whole, the method is quite promising and will be the subject of further study by me, which I trust may result in a somewhat more accurate and more convenient method than those now in general use.

AGRICULTURAL EXPERIMENT STATION,
MORGANTOWN, WEST VIRGINIA,
December 10, 1894.

ON CERTAIN PHENOMENA OBSERVED IN THE PRECIPITATION OF ANTIMONY FROM SOLUTIONS OF POTASSIUM ANTIMONYL TARTRATE.

BY J. H. LONG.

Received December 19, 1894.

I HAVE elsewhere called attention to the behavior of solutions of tartar emetic when treated with solutions of other salts, (see *Am. Jour. Sci. and Arts*, October, *1889* and October, *1890*) and with Mr. H. E. Sauer have determined the conditions of precipitation by carbonates, acetates, and phosphates. (*J. Anal. Appl. Chem.*, March, *1891*.)

When to solutions of the antimony salt sulphates, chlorides, nitrates, and oxalates of the alkali metals are added no precipitation occurs, even with elevation of temperature. With carbonates, acetates, phosphates, borates, thiosulphates, sulphites, tungstates, and some other compounds, clear solutions can be made at a low temperature, but precipitation follows at a higher point. The precipitate, in nearly all cases, consists of hydrated antimony oxide, and its amount is a function of time, temperature, and amount of added salt.

With mixtures of the tartrate and sodium carbonate, for instance, it was found that in the cold, at the end of twenty-four hours, amounts were precipitated as shown in the following table. In each test five grams of the tartrate were dissolved in sixty cc. of warm water and cooled to 20°. Then different weights of pure sodium carbonate were dissolved in thirty-five cc. of water; these solutions were added to the others and the mixtures were brought up to 100 cc. They were allowed to stand until precipitation was complete, usually over night or longer. An aliquot part of the clear supernatant liquid was

taken and the amount of antimony in solution determined. This was calculated to tartrate in the whole:

No. of experiment.	Na_2CO_3 added.	$KSbOC_4H_4O_6.\frac{1}{2}H_2O$ left in solution.
1	0.1 gram	99.93 per cent.
2	0.3 "	85.22 " "
3	0.5 "	70.36 " "
4	0.7 "	56.76 " "
5	0.9 "	40.87 " "
6	1.2 grams	29.17 " "
7	1.5 "	13.94 " "
8	2.0 "	3.88 " "

In another series of experiments the solutions of carbonate and tartrate were mixed as before at 20° and then brought to 100°, and maintained at this heat one hour. The precipitates formed immediately, and at the end of the hour were separated by filtration. The filtrates were tested for antimony remaining. The results are shown in the third column below.

No. of experiment.	Na_2CO_3 added.	$KSbOC_4H_4O_6.\frac{1}{2}H_2O$ left in solution.
9	0.2 gram	79.23 per cent.
10	0.5 "	46.70 " "
11	0.8 "	21.74 " "
12	1.1 grams	8.69 " "
13	1.5 "	6.33 " "
14	2.0 "	4.42 " "
15	3.5 "	4.66 " "
16	5.0 "	4.74 " "

On comparing the two tables, it will be seen that at first the precipitation is much more rapid in hot solution than in cold, but that finally, with excess of carbonate, a more complete decomposition of the tartrate is effected in the cold solution. Two equations can be given, according to which the reaction may take place. The first of these is:

$$2KSbOC_4H_4O_6 + 2Na_2CO_3 + 2H_2O = 2KNaC_4H_4O_6 + Sb_2O_3.H_2O + 2HNaCO_3.$$

The second is:

$$2KSbOC_4H_4O_6 + Na_2CO_3 + H_2O = 2KNaC_4H_4O_6 + Sb_2O_3.H_2O + CO_2.$$

The first, probably, takes place in the cold solution as no carbon dioxide escapes. The loss of carbon dioxide from the hot

solution is less than called for by the equation because an excess of neutral carbonate is present and the solutions are not actually boiled. In any case the precipitation is incomplete, and by addition of increased amounts of sodium carbonate, a condition is reached in which a part of the oxide at first thrown down appears to go into solution again.

Precipitation with sodium acetate takes place imperfectly in the cold, but by heat a stronger reaction follows. In both cases it was found that the results may be expressed by the following equation:

$$2KSbOC_4H_4O_6 + 2NaC_2H_3O_2 + 2H_2O = 2KNaC_4H_4O_6 + Sb_2O_3.H_2O + 2HC_2H_3O_2.$$

With phosphates the experiments led to the conclusion that precipitation takes place in a manner represented by this equation:

$$2KSbOC_4H_4O_6 + HNa_2PO_4 + 2H_2O = 2KNaC_4H_4O_6 + Sb_2O_3.H_2O + H_3PO_4.$$

With cold solutions precipitation is very slow, but by heat an amount of the antimony oxide corresponding to seventy-five per cent. of the tartrate originally in solution is obtained.

I have since investigated the behavior of several other salts as precipitants, with the results which follow.

REACTION WITH SODIUM BIBORATE.

A very sharp reaction takes place between solutions of borax and tartar emetic, which was studied in the following manner: In the first series of experiments, five grams of the tartrate were dissolved for each test in sixty cc. of water, the solutions being brought to 20°. To these were added definite weights of borax dissolved in thirty to thirty-five cc. of water at the same temperature. The mixtures were made up to 100 cc. exactly, and allowed to stand over night in a place with nearly constant temperature. In all cases a precipitate formed which was separated by filtration. The analysis of the precipitate showed it to have the same composition as that formed by the sodium carbonate; *viz.*, Sb_2O_3.aq. On drying at a high temperature most of the water is lost, leaving practically Sb_2O_3.

In each case the precipitate was separated by filtration, and

the filtrate made up to 250 cc. Twenty-five cc. of this was taken and precipitated by hydrogen sulphide, after addition of tartaric and hydrochloric acids in small amount. The precipitation was finished on a hot water-bath, and the precipitate collected on a Gooch filter, washed, dried at 120°, and weighed. The sulphide was calculated to tartrate on the supposition that all the antimony in solution was left in the original form (Sb=120, O=16). The results obtained are shown by these figures:

No. of experiment.	Borax added.	$KSbOC_4H_4O_6.\frac{1}{2}H_2O$ left in solution.
1	0.1 gram	98.81 per cent.
2	0.2 "	94.74 " "
3	0.4 "	87.78 " "
4	0.8 "	74.98 " "
5	1.6 grams	46.84 " "
6	3.2 "	3.31 " "

In another series of tests the solutions containing the borax and tartrate were made up to 250 cc. instead of to 100 cc. They were allowed to stand, filtered, and treated as before, giving these results:

No. of experiment.	Borax added.	$KSbOC_4H_4O_6.\frac{1}{2}H_2O$ left in solution.
7	0.1 gram	100.00 per cent.
8	0.2 "	97.15 " "
9	0.4 "	89.05 " "
10	0.8 "	75.29 " "
11	1.6 grams	49.09 " "
12	3.2 "	6.01 " "
13	6.4 "	2.03 " "
14	12.8 "	0.99 " "

In a third set of experiments the liquid containing the borax and tartrate was diluted to 100 cc. in a flask, as in the first set. The flask was closed with a perforated rubber stopper having a long glass tube attached, and then heated in boiling water one hour. The liquid was allowed to cool, was filtered, and the filtrate made up to 250 cc. An aliquot part, on analysis, gave results which are shown below.

No. of experiment.	Borax added.	$KSbOC_4H_4O_6.\frac{1}{2}H_2O$ left in solution.
15	0.1 gram	99.88 per cent.
16	0.2 "	95.37 " "
17	0.4 "	88.81 " "
18	0.8 "	74.78 " "
19	1.6 grams	47.03 " "
20	3.2 "	3.96 " "
21	6.4 "	1.76 " "

In these tables several things are immediately apparent. It appears that the precipitation is less perfect in hot solution than in cold, although for equal weights of borax the differences are not great. It is seen also that the amounts precipitated are greater in the strongest solutions. These effects of temperature and concentration are far less marked, however, with borax precipitation than with that by the sodium carbonate.

With borax we have, in each case, a very regular rate of precipitation. By platting the weights of borax in the above table as abscissas, and the amounts of tartrate left as ordinates, we obtain a curve which is almost a straight line.

It seems practically impossible to precipitate all the antimony by excess of borax, although the amount left in solution is much less than when sodium carbonate was used as the precipitant. Direct trials showed that the solubility of the antimonous oxide in excess of borax solution is very slight, but is a trifle greater in the excess of sodium carbonate. The solubility in the Rochelle salt solution formed in the latter case will not account for this difference as will appear below.

In order to gain further insight into the reaction I measured the amount of rotation of polarized light in a number of solutions before and after the separation of the precipitate of antimonous oxide. Some exceedingly interesting results were obtained, a few of which will be explained in detail. I dissolved five grams of the tartrate in fifty cc. of hot water and added three grams of borax in twenty-five cc. of water, made up to ninety cc., and heated one-half hour in the water-bath. The solution was allowed to cool to 20° and made up to 100.5 cc. (on account of volume of precipitate), and filtered through a dry filter. The filtrate was polarized in a 200 mm. tube, giving

$$\alpha_D = 3.596°.$$

Seventy-five cc. of the filtrate, after the addition of a little hydrochloric and tartaric acids, was precipitated by hydrogen sulphide. The precipitate was collected, washed, and dried in the usual manner in the Gooch funnel. I found 0.119 gram of the sulphide, corresponding to 0.312 gram of $KSbOC_4H_4O_6.\frac{1}{2}H_2O$, in the whole filtrate. 4.688 grams had, therefore, been precipitated. From the outset it would naturally occur to one that the

precipitation of antimonous oxide must be accompanied by the formation of sodium potassium tartrate, and that the polarization effect observed must, in part, be due to this, as well as to the potassium antimonyl tartrate left. I have elsewhere given the results of very accurate tests in which the rotation constants of these tartrates were determined by the use of the large Landolt-Lippich polarimeter with the 400 mm. tube (*Am. Jour. Sci. and Arts, loc. cit.*). From these it appears that the rotation of 0.312 gram of the $KSbOC_4H_4H_4O_6.\frac{1}{2}H_2O$, and 3.982 grams of $KNaC_4H_4O_6.4H_2O$ (this latter corresponding to the tartar emetic decomposed) in 100 cc. should not be over 2.6°. In the direct polarization of the filtrate I found, as given above, 3.596°. It is evident, therefore, that something else must be present to modify the result.

It is well known that the presence of boric acid increases the rotation of tartrates in a marked degree, and this can be readily accounted for here if we assume that the reaction takes place according to the following equation:

$$2KSbOC_4H_4O_6 + Na_2B_4O_7 + 6H_2O = 2KNaC_4H_4O_6 + Sb_2O_3 + 4H_3BO_3.$$

On applying tests for free boric acid its presence was readily shown. We have here apparently a reaction similar to those in which acetic and phosphoric acids are liberated from acetates and phosphates.

In the last experiment it was shown that antimony, corresponding to 0.312 gram of the potassium antimonyl tartrate was still in solution, or that 4.688 grams had been decomposed. To do this according to the above equation would require 2.697 grams of crystallized borax, and would leave in solution 3.982 grams of $KNaC_4H_4O_6.4H_2O$ and 1.751 grams of H_3BO_3. An excess of 0.303 gram of borax would be left in solution. To test the correctness of this view I prepared a solution containing in 100 cc., at 20°:

0.312 gram of $KSbOC_4H_4O_6.\frac{1}{2}H_2O$.
3.982 grams of $KNaC_4H_4O_6.4H_2O$.
1.751 grams of H_3BO_3.
0.303 gram of $Na_2B_4O_7.10H_2O$.

This solution was polarized in the 200 mm. tube and gave

$$\alpha_D = 3.590°,$$

which agrees very well with the result of the first experiment. Another solution, containing in 100 cc., at 20°,

0.150 gram of $KSbOC_4H_4O_6.\frac{1}{2}H_2O$,
4.119 grams of $KNaC_4H_4O_6.4H_2O$,
1.811 grams of H_3BO_3,

gave $\alpha_D = 3.661°$. While boric acid increases the rotation of tartrates and tartaric acid I have elsewhere shown that borax decreases the rotation of Rochelle salt slightly. The equation probably represents the facts properly.

In the above nothing has been said about certain peculiarities observed in the formation of the precipitates. When cold dilute solutions of borax and the tartrate are mixed no reaction takes place immediately, but with warm, strong solutions, a precipitate seems to form as soon as the two liquids are poured together. In a former paper (*J. Anal. Appl. Chem.*, *loc. cit.*) I pointed out the important and exceedingly curious fact that in the reaction between carbonates or acetates on the one hand, with the tartrate on the other, while no precipitate may appear immediately, perhaps not in hours, indicating a decomposition, the polarimeter shows that such has taken place. Here, also, we have evidence that a reaction has taken place even without precipitation, and this the polarimeter furnishes. The matter can be best explained by giving the details of several experiments.

I made five solutions by dissolving five grams of the tartrate as before in fifty-five to sixty cc. of warm water, cooled to 20°, and added certain weights of borax in small volumes of water, making the solutions finally to 100 cc. at 20°. These solutions were polarized immediately in the 200 mm. tube, with the following results:

No. of experiment.	$KSbOC_4H_4O_6.\frac{1}{2}H_2O$ taken.	$Na_2B_4O_7.10H_2O$ added.	α_D
1	5 grams	0.5 gram	12.08°
2	5 "	1.0 "	9.40°
3	5 "	2.0 grams	5.53°
4	5 "	3.0 "	4.10°
5	5 "	4.0 "	3.75°

The normal rotation of the tartrate at 20° in the 200 mm. tube, with a concentration of five grams in 100 cc., I have shown to be

$$\alpha_D = 14.103°.$$

The effect of the borax is therefore marked, but the extent of the decrease in rotation depends on the number of minutes intervening between the mixing of the solutions and the completion of the observation in the polarimeter. A gradual decrease in the readings was in all cases observed, until the solutions became finally too turbid for observation from the beginning precipitation. The first solution, for instance, in the above table was read as follows:

10	hours,	30	minutes,	$\alpha_D = 12.08°$
10	"	45	"	" $= 11.87°$
10	"	55	"	" $= 11.80°$

After standing some hours, the solutions deposited a precipitate and cleared up. On again polarizing I found:

No.	α_D
1	11.52°
2	9.01°
3	5.29°
4	3.92°
5	3.65°

These observations were made in a 200 mm. tube, but similar solutions were polarized in a 400 mm. tube with perfect sharpness, the readings agreeing within 0.02°, as is possible with the large and excellent instrument used. I mention this to prove the perfect transparency of the liquids, and to show that the decreased rotations observed at the start were not due to any loss through precipitation, but were in consequence of changes preceding precipitation, these changes taking place very gradually.

We have here a phenomenon reminding one of the birotation of solutions of certain sugars, but depending on a different cause undoubtedly. In the reaction between the same tartrate and sodium carbonate the same change was observed, but through a longer period. A solution containing in 100 cc. one-tenth gram of the carbonate and five grams of the tartrate gave, at the end of five minutes in a 400 mm. tube, a rotation of 25.582°, after

thirty minutes, 25.580°, that is, practically the same, but after twelve hours, 24.480°. A perfectly clear mixture can be made containing five grams of the tartrate and nine-tenths gram of sodium carbonate in 100 cc. When polarized immediately, I found with this in the 400 mm. tube $\alpha_D = 11.57°$, after ten minutes, 11.50°, after twenty-five minutes, 11.132°, and after sixty-five minutes, 10.55°.

In the normal reaction between carbonates or borates and the potassium antimonyl tartrate a precipitate should be formed, but we find that at a low temperature this is much delayed. If precipitation alone were taken as the indication of a reaction it would necessarily appear that at the outset no reaction takes place, but the behavior with polarized light shows the error in this view. It is evident that a reaction begins immediately and progresses far toward completion in some of the cases considered before even the first polarization can be made, that is within two or three minutes. This first part of the reaction is the beginning stage of precipitation and may consist in the formation of some intermediate product, which finally decomposes. I have elsewhere shown (*Am. Jour. Sci. and Arts, loc. cit.*) that the rotation of potassium sodium tartrate is decreased by the addition of sodium, thallium, and lithium salts, but is increased by the addition of potassium and ammonium salts and that this reaction is fully accounted for if we assume the formation of sodium tartrate, sodium thallium tartrate or sodium lithium tartrate in the one case or of potassium or potassium ammonium tartrate in the other. In the present instance we evidently must admit the formation of sodium potassium tartrate from the instant the solutions are mixed, but that the reaction is a progressive one. The potassium antimonyl compound with a high rotation gives place to the potassium sodium compound with a much lower rotation. There is nothing to show, however, in what form the antimony is held.

Possibly the reaction may be explained by assuming the formation of an intermediate product according to this equation:

$$2KSbOC_4H_4O_6 + Na_2B_4O_7 = 2KNaC_4H_4O_6 + (SbO)_2B_2O_7.$$

If the last compound is formed it must break up in this manner:

$$(SbO)_2B_4O_7 + 6H_2O + XH_2O = 4H_2BO_3 + Sb_2O_3.XH_2O,$$

leaving a hydrated oxide of antimony with more water than the final precipitate contains. By loss of water, possibly, this hydrated compound must, in time, settle out as a precipitate. It has been explained that by boiling, the precipitate forms and subsides soon. At 20°, even after what I have called the preliminary stage of precipitation, may have occupied hours, the actual formation of the precipitate may consume an equally long time. The precipitate is a growth through an invisible and a visible stage, and what is true here is true of the next case to be given.

REACTION WITH SODIUM TUNGSTATE.

Cold solutions of the tartrate give no immediate precipitate when mixed with cold solutions of ordinary sodium tungstate, but on standing the mixtures gradually become turbid and finally deposit a sediment. The composition of this depends largely on the temperature and concentration. The precipitate formed in the cold, collected and dried at 105°–110° C. consists, essentially of antimony oxide. A precipitate formed by mixing hot solutions has practically the same composition, but if obtained after long heating it contains a relatively larger amount of tungstic acid.

In a series of tests made by mixing hot solutions of the two salts, and allowing the mixtures to stand several hours to cool, the following results were obtained :

$KSbOC_4H_4O_6.\frac{1}{2}H_2O$ in 100 cc.	$Na_2WO_4.2H_2O$ in 50 cc.	Wt. of ppt.	Sb_2S_3 from same.	Per cent. of Sb. in ppt.
1 gram	2 grams	0.3087	0.3467	80.22
2 grams	2 "	0.6442	0.7291	80.84
4 "	2 "	0.5728	0.6486	80.88
5 "	2 "	0.7222	0.8170	80.79

The mixtures were made in platinum dishes holding about 200 cc., and as the precipitates formed as a coherent coating on the dishes they were easily washed, dried at 110°, and weighed. They were then dissolved in diluted hydrochloric acid, which left a small amount of tungstic acid in each case in flocculent form. The solutions were then filtered, and, after the addition of some tartaric acid, were precipitated by hydrogen sulphide in the usual manner. The sulphide precipitates were collected on a Gooch, dried at 110°, and weighed. It will be seen that the

results are a little low to correspond to pure antimony oxide as the composition of the white precipitate. This compound contains 83.3 per cent. of antimony. The compound $Sb_2O_3.H_2O$ contains 78.4 per cent. The lower results are doubtless due to the small amounts of tungstic acid left in each case on treatment with hydrochloric acid, and referred to above.

In a second series of experiments constant amounts of the tartrate in hot solution were mixed with varying amounts of the tungstate likewise in hot solution. The white precipitates which formed were collected and weighed as before, with the following results, which show the effect of excess of tungstate on the amount of precipitate. In each case two grams of the tartrate was taken in 100 cc., and the tungstate in fifty cc.:

Tungstate taken.	Ppt. obtained.
0.5 gram	0.1070
1.0 "	0.4598
2.0 grams	0.5762
3.0 "	0.5861
4.0 "	0.6143
5.0 "	0.6185

From this, it is plain that the amount of precipitate is not much increased by great excess of the tungstate beyond a certain point. In the cold, precipitation is much less perfect, while, by boiling, fully three-fourths of the theoretical yield of oxide from the tartrate can be obtained.

By working with cold solutions a mixture may be made which remains clear long enough to permit polarimeter observations to be taken. I dissolved five grams of the tartrate in sixty cc. of water, cooled to 20°, and added five-tenths gram of the tungstate in twenty cc. of water. The mixture was made up quickly to 100 cc. at 20° and polarized in the 200 mm. tube immediately and after intervals of five minutes. I found without refilling the tube:

1st observation	α_D	$= 11.66°$
2nd "	"	$= 12.03°$
3rd "	"	$= 12.13°$
4th "	"	$= 12.53°$

The solution became now too turbid, from separation of a precipitate, for further tests. On standing some hours the remainder

of the solution in the flask cleared after subsidence of its precipitate. A portion of this examined gave

$$\alpha_D = 12.74^\circ.$$

Another portion of the same solution, heated and then cooled to 20°, gave the same. A new solution prepared in the same manner gave, after standing some time,

$$\alpha_D = 13.11^\circ.$$

I made next a solution containing in 100 cc. five grams of the tartrate and one gram of the tungstate. This examined, immediately, at 20°, gave

$$\alpha_D = 9.42^\circ,$$

but it soon became turbid and deposited a precipitate. After clearing, I found

$$\alpha_D = 11.48^\circ,$$

which increased to 11.87° by heating the liquid.

These reactions are especially interesting when compared with those between the tartrates and other salts. We have here, as before, a marked decrease in the specific rotation on mixing the solutions of the active and inactive substances. But in the case of the tungstate, on standing, there is an increase instead of a decrease in the rotation observed in the other cases. This behavior finds its explanation probably in the action of the liberated tungstic acid. The reaction between the two salts undoubtedly follows this equation:

$$2KSbOC_4H_4O_6 + Na_2WO_4 = 2KNaC_4H_4O_6 + Sb_2O_3 + (WO_3)x.$$

The precipitation of the tungstic acid is very slow and incomplete. While in solution, it may combine with the soluble tartrate to form a body with increased rotation, the possibility of which is shown by the researches of Gernez and others. The delay in the appearance of the precipitate may be due to the formation and slow breaking up of intermediate products containing the antimony and tungstic oxides in temporarily soluble condition. With liberation of the tungstic acid we have a gradual increase in the already decreased rotation. This change in the rotation, before precipitation, is well shown in the following observations. I dissolved five grams of the tartrate and two and five-tenths grams of the tungstate, mixed at a low tempera-

ture and made up to 100 cc. as before, and at exactly 20° C. A reading with the 200 mm. tube was made as soon as possible and, without changing the solution, this was repeated at frequent intervals. The results were as follows:

3 hours, 28 min.	$\alpha_D = 5.66°$
	" $= 5.96°$
	" $= 6.45°$
	" $= 6.83°$
	" $= 7.15°$
	" $= 7.32°$
3 hours, 44 min.	" $= 7.43°$
3 hours, 58 min.	" $= 7.90°$

At this point the liquid began to grow slightly turbid, so that the observations could not be continued. The remainder of the liquid was then heated in a closed flask to complete the precipitation, cooled to 20°, and tested. I found now

$$\alpha_D = 9.13°.$$

The liquid still remaining was filtered, fifty cc. of the filtrate taken and precipitated with hydrogen sulphide, yielding finally 0.4045 gram of antimony sulphide. From this it appears that of the tartrate originally taken, 1.599 grams remained in solution in the 100 cc. (no allowance being made for the volume of the precipitate).

From this we have

$KSbOC_4H_4O_6.\frac{1}{2}H_2O$	decomposed =	3.401 grams.
" "	remaining	1.599 "
$KNaC_4H_4O_6.4H_2O$	formed	2.888 "

Now, the rotation corresponding to the tartar emetic remaining is 4.50° and that to the Rochelle salt formed is 1.25° from which we should expect a total rotation of 5.75°. This, in fact, corresponds very nearly to what was observed at the beginning of the test, and seems to bear out the suggestion made above: *viz.*, that the principal reaction here occurs before actual precipitation appears. A splitting of the tartar emetic is indicated by the immediate decrease in the rotation and then the complex effect of addition of the liberated tungstic acid to the alkali tartrate in solution appears from the gradually increased rotation

Precipitation finally follows as the end of the reaction; the separated part assumes the insoluble form.

It will be recognized that the phenomenon in this case is much more complex than in the other. There is nothing to show that while the rotation is being increased by the action of the tungstic acid there is not also a tendency toward decrease because of progressive decomposition of the potassium antimonyl tartrate. In all probability the observed rotation is a resultant effect.

REACTION WITH SODIUM THIOSULPHATE.

Cold dilute solutions of the thiosulphate and potassium antimonyl tartrate can be mixed without immediate precipitation. Application of heat, however, produces a light yellow precipitate which grows deeper and finally becomes bright red. This precipitate is the substance commonly known as antimony cinnabar, used as a pigment, and on the large scale is made by decomposing the thiosulphate by antimony chloride. In the reaction between the thiosulphate and tartar emetic the precipitate appears very heavy, but the decomposition is far from complete as shown by the figures given below.

There seems to be some doubt as to the composition of this precipitate. Roscoe and Schorlemmer (**2**, part 2, 324) give it as, probably, SbS_2O, referring, however, to other formulas. Dammer's Handbuch gives $Sb_2O_3.Sb_2S_3$ as the probable formula. Others are also given. Recently Baubigny (*Compt. rend.*, No. 17, *1894*), has given reasons why the formula Sb_2S_3 should be considered the correct one and the proof he presents appears to be satisfactory. There remains a possibility, however, that the composition may, under certain circumstances, vary with the method of preparation. In fact, some of my own results seem to show this and I am now engaged in studying the question further. But as made in the reaction in hand the composition seems to be $2Sb_2S_3.H_2O$. This was determined by the following considerations: The precipitate dissolves in hydrochloric acid without liberation of sulphur, yielding a perfectly clear solution. A solution made in this way was heated, mixed with a little tartaric acid solution, and precipitated by hydrogen sulphide in the usual manner. On filtering off the orange-yellow precipi-

tate so obtained in a Gooch crucible, drying at 120°, and weighing, the weight was always found less than that of the antimony cinnabar taken.

Sulphur determinations were made by dissolving a gram or less of the substance in strong S-free solution of potassium hydroxide and then oxidizing the sulpho-salt formed by washed chlorine gas, (method of Rivot). The results of these tests were as follows:

	Calculated for $2Sb_2S_3.H_2O$.	Found.
Sb	69.56	69.80
S	27.83	27.72

No determination of the water was made, but its presence was shown in the substance dried at 120° by heating to a higher temperature in a narrow glass tube.

In a series of experiments on the precipitation a number of portions of the tartrate, of five grams each, were weighed out and dissolved in 150 cc. of water. Varying amounts of the thiosulphate in fifty cc. of water were added and then water enough to make exactly 250 cc. The flasks holding the mixtures were closed with perforated stoppers containing long glass tubes and then heated in the water-bath one hour. In this way evaporation was practically avoided. At the end of the hour the precipitates were collected on a Gooch funnel, dried at 120°, and weighed with the following results:

No. of experiment.	Weight of $Na_2S_2O_3.5H_2O$ added.	Weight of precipitate.
1	0.1 gram.	0.0039 gram.
2	0.2 "	0.0068 "
3	0.4 "	0.0111 "
4	0.8 "	0.0178 "
5	1.6 grams.	0.0363 "
6	3.2 "	0.0803 "
7	6.4 "	0.2192 "
8	12.8 "	0.4809 "

In a second set of tests I dissolved, in each case, ten grams of the tartar emetic in 100 cc. of hot water and added the thiosulphate in fifty cc. of hot water. The mixtures were kept at 100° two hours and then filtered. Results as follows:

No. of experiment.	Weight of $Na_2S_2O_3.5H_2O$ added.	Weight of precipitate.
9	5 grams.	0.1915
10	10 "	0.4041
11	15 "	0.5162
12	20 "	0.6818

It is evident from these figures that in both sets of experiments the reaction is far from complete and not easily determined. It is, perhaps, quite complex. I noticed in no case the escape of hydrogen sulphide or sulphurous oxide and the gradual change in color during precipitation from very light yellow to bright red suggests that it takes place in two stages. Vortmann (*Ber. d. chem. Ges.*, **22**, 2307) has studied the general problem of decomposition of thiosulphuric acid and states that it breaks up into H_2S, O, and SO_2. In presence of certain metals, tetra- and pentathionates seem to be formed. It is certain that no sulphate is formed in the reaction in hand, but the proof of formation of the several thionic acids is difficult because of the incompleteness of the reaction and the presence of the great excess of thiosulphate. In cases of complete reaction, however, Vortmann and Vaubel, also, (*Ber. d. chem. Ges.*, **22**, 2703) have shown that these acids are formed.

I suggest, therefore, this explanation of the present reaction. At the outset there may be, as with borax, a decomposition according to this equation,

$$2KSbOC_4H_4O_6 + Na_2S_2O_3 + H_2O = Sb_2O_3 + 2KNaC_4H_4O_6 + H_2S_2O_3,$$

then

$$Sb_2O_3 + 2H_2S_2O_3 = Sb_2S_3 + 2H_2O + SO_2 + O_3,$$

the oxygen and sulphur dioxide being held, however, to form polythionates.

The gradual change of color can be accounted for by the gradual change of oxide into sulphide of antimony and it seems possible that under certain conditions of concentration and temperature a part of the oxide should remain unchanged, accounting for the results of some of the analysts who have examined the precipitate. In several instances I modified the experiment by mixing warm solutions of the tartar emetic with warm thiosulphate solutions, and then throwing the mixture into a large

volume of cold water the instant a precipitate began to form, to check the reaction. The precipitate which now settled was very light colored instead of red, and appeared to be a mixture of oxide and sulphide. The microscopic appearance of this precipitate is distinctly different from that of the antimony cinnabar.

As explained at the outset some time elapses after mixing cold solutions before a precipitate appears. In the beginning of the interval the solution may be so clear that accurate polarimetric observations are possible. But, contrary to expectations, based on the behavior in other cases recorded, I find the specific rotation of the tartar emetic, in this stage of the reaction, quite unchanged. After precipitation a marked decrease was observed as usual. This is shown in the following table in which the filtrates from the last precipitates referred to were made up to 250 cc. before polarization:

No. of experiment.	Thiosulphate added.	α_D observed.	α_D calculated.
9	5 grams.	10.90°	10.89°
10	10 "	10.43°	10.46°
11	15 "	10.07°	10.24°
12	20 "	9.69°	9.93°

In the last column the rotations were calculated on the assumption that antimony is precipitated and Rochelle salt formed according to the last equations. In the eleventh and twelfth experiments the thiosulphate is present in amount much in excess of that necessary to complete precipitation by these reactions, and it is evident that a decomposition of the active molecule has taken place not indicated by the amount of precipitate.

There seems to be a fundamental difference between the reaction with the thiosulphate on the one hand, and those with the borate, carbonate, phosphate, acetate, and tungstate on the other. In the first case we appear to have no breaking up of the complex tartrate until actual precipitation occurs, while in the others the stages are quite distinct. I have shown that in these last reactions acids are liberated which may be readily recognized. It is also true that these acids are practically without action on antimonous oxide, from which it would follow that this substance might remain a long time, possibly in the

hydrated form, in contact with the liberated acids without change. This would not be the case with liberated thiosulphuric acid. If set free in the presence of antimonous oxide, even the dry precipitate, it soon converts it into sulphide. Any cause, therefore, which acts to destroy the equilibrium in the solution of tartrate and thiosulphate must lead to the almost immediate formation of a precipitate.

In many of our most familiar cases of precipitation the formation of the precipitate is so rapid that we are accustomed to look upon it as instantaneous. But by varying the conditions of precipitation it may be shown that even the reaction between barium chloride and sulphuric acid is one which consumes an appreciable interval. In such cases, however, we have no means of knowing what takes place before the precipitate becomes actually visible. It is, doubtless, true that the liquid regarded as supersaturated for $BaSO_4$ does not begin to throw this out in solid form until a relatively large number of these molecules combine to produce a particle of a certain size, but at what rate the Ba and SO_4 ions combine can not be shown.

But in the cases before us we have evidence, first, of the gradual breaking up of the complex potassium antimonyl tartrate and then, also, that a relatively large amount of the antimony may be separated before any of it falls as a precipitate. The stage of precipitation seems to follow as a perfectly distinct and also progressive one. It is hastened, as in other cases of supersaturation, by heat or agitation. One of the reactions shows, also, that a relatively long time is consumed in combination as well as in dissociation. In the case of the mixture containing the tartrate and tungstate we have evidence of the splitting of the first molecule, and then, from the slowly increasing rotation, evidence of the addition of an element from the second to the active part of the first. Both of these phenomena are observed before precipitation appears.

It must be remembered that the several acids shown to be liberated in these experiments are all so-called weak acids, or acids but slightly dissociated in solutions. It is also true, as a test of their electric conductivities shows, that the solutions of tartar emetic are relatively little dissociated (see in this connec-

tion Hugo Haedrich, *Ztschr. phys. Chem.*, **12**, 496). There is doubtless, therefore, a close connection between the phenomena outlined above, and others depending on the degree of dissociation of the ions in solution. A study of the behavior of weaker solutions would doubtless lead to interesting results not brought out in the above experiments which were undertaken mainly to show the character and amount of precipitates formed. A discussion of the behavior of dilute solutions will follow.

CHICAGO, December, 1894.

[CONTRIBUTIONS FROM THE CHEMICAL LABORATORY OF CASE SCHOOL OF APPLIED SCIENCE. XXI.]

AN EXAMINATION OF THE ATMOSPHERE OF A LARGE MANUFACTURING CITY.[1]

BY CHARLES F. MABERY.

Received December 13, 1894.

THE composition of the atmosphere was determined in the latter part of the last century by the classic researches of Priestley, Cavendish, Lavoisier, and others. These investigations were stimulated by Priestley's discovery of oxygen and the discovery of nitrogen by Rutherford. The proportion of nitrogen was first ascertained by Cavendish, and the precise quantitative composition was ascertained by Lavoisier, who repeated the celebrated experiment of Priestley, in which metallic mercury was heated for a long time in a confined volume of air to determine the quantity of gas absorbed and the residual nitrogen. The composition of the atmosphere with reference to the proportion of oxygen and nitrogen has since been frequently determined without showing any appreciable variation. In specimens of air collected from regions widely separated, Morley (*Am. Chem. J.*, **2**, 276) determined with the utmost precision the proportion of oxygen and nitrogen without finding any appreciable variation.

The essential constituents of the atmosphere include oxygen, nitrogen, and carbon dioxide in constant proportions, nitrous and nitric acids, sulphurous and sulphuric acids, dust, soot, moisture, ammonia, and certain other constituents in variable

[1] The experimental work of this paper was performed under my direction by Mr. Charles G. Snyder, and it formed the subject of his thesis for the Degree of Bachelor of Science.

proportions depending upon atmospheric conditions and local influences. The allotropic form of oxygen, ozone, is an important constituent of the atmosphere, and probably also hydrogen dioxide.

With a population distributed over large areas, natural conditions are sufficient to keep the atmosphere in a normal and healthful condition for the support of life. Such has been the condition in all civilized countries until within the last thirty years the rapid accumulation of population near manufacturing centers and in cities, and the enormous consumption and incidental waste of coal has introduced important changes at these points. Not only are manufactories and other large establishments dependent for their maintenance upon the consumption of coal, but private habitations, which were formerly heated with wood as fuel with the production of a harmless smoke, now depend almost entirely upon coal. If hard coal is burned the smoke is not objectionable, but the escape of bituminous coal smoke from every household chimney in a densely populated city, together with dense volumes poured forth from large blocks is sufficient to keep the atmosphere in a questionable condition in relation to health, comfort, and the deleterious effects on perishable forms of property. Under occasional conditions of strong winds or rain, followed by clear weather, the city atmosphere is doubtless in a harmless condition; but it can not long remain unpolluted with immense volumes of sooty smoke poured forth from many chimneys.

Abroad, especially in England, the atmosphere of large cities, notably in manufacturing towns, has received considerable attention. Doubtless, on account of the rapid growth of American cities, and probably, also, because of an abundant supply of anthracite coal for cities too far removed from the regions of bituminous coal to render it profitable, the atmosphere of our cities has received little attention. But for cities situated near the soft coal regions, the question of atmospheric pollution has become of too serious moment to be long disregarded.

To understand the nature of smoke pollution it is necessary to refer to the composition of coal and to the products of its combustion under different conditions. Bituminous coal consists

principally of compounds of carbon and hydrogen, or hydrocarbons with smaller proportions of nitrogen and sulphur compounds, and metallic compounds known as the ash when the coal is burned. So far as atmospheric pollution by coal smoke is concerned, the four elements, carbon, hydrogen, sulphur, and nitrogen, are the only constituents of coal that need to be considered.

In the combustion of coal with an adequate supply of air, under such conditions that the chemical reactions involved can be completed within the space allotted to combustion, the ultimate products are carbon dioxide, carbon monoxide, water, sulphurous acid, and ammonia. The less carbon monoxide the smoke contains the more efficient is the combustion; the sulphurous acid is usually converted, at least in part, into sulphuric acid by further oxidation in presence of moisture. It is only rarely, however, that the conditions requisite for complete combustion are maintained. The following features will be recognized as having an important bearing upon the question of complete combustion: (1) A sufficient supply of air; (2) the combustion should not be forced too rapidly; (3) the temperature, especially at the moment of firing, must not be too low; (4) the size, form, and proportions of the space within which the combustion takes place must be suitably adapted. These are the principal points with which complete combustion is concerned, and they are evidently closely connected with the manner of firing and general care of the heating apparatus. In practice, these conditions are far from being realized; with an insufficient supply of air unequally distributed with reference to the fuel, especially with a low temperature, the carbon is only partially burned and the portion not consumed escapes in the form of soot.

The hydrogen burns first, and it has an important influence in maintaining the temperature sufficiently high for complete combustion of the carbon. If the coal contain considerable sulphur it must escape as sulphurous acid, or as sulphuric acid in complete combustion, or (as hydrogen sulphide) in incomplete combustion, and it would seem to be impossible to avoid deleterious effects of these acids in the atmosphere. No doubt sul-

phuric acid should have some effect on such materials as bookbindings and delicate fabrics after collecting in dust during long periods of time, but the effect is greatly increased by the presence of soot, which, as we shall show, absorbs the acid, and when it collects on any surface it occasions a greater accumulation of the acid than should occur in its absence.

The state of the atmosphere in cities where bituminous coal is the principal fuel is dependent upon conditions of combustion in which the chemical reactions are far from complete, with the result that the carbon escapes from the chimney in part as carbon dioxide, in part as carbon monoxide, and in a considerable part as soot. Under such conditions the sulphur is set free as a mixture of sulphuric and sulphurous acids and hydrogen sulphide. The latter is readily shown by its effects on metals, especially on silver, which soon becomes badly discolored on account of the formation of argentic sulphide. The soot ordinarily present in the atmosphere has a greasy consistency, readily recognized by subjecting it to pressure. It consists to a considerable extent of compounds of carbon and hydrogen known as hydrocarbons, which escape from the coal by a process of destructive distillation under conditions somewhat similar to those in the distillation of coal in gas retorts. As mentioned above it carries with it the sulphur acids and ammonia, and it is highly destructive to vegetable life, its action depending probably in part upon the formation on the leaves and stems of a thin layer of tarry substance, which clogs the pores, thus interfering with the process of plant respiration. Sulphuric acid is highly destructive to the forms of organic compounds of which vegetation is composed and in soot it may accumulate to such an extent as to be injurious.

In complete combustion with no escape of soot, even on a very large scale, as in manufactories, it is safe to assert, that the natural conditions of the atmosphere may prevent any undesirable effect from the products given off.

During the year 1889 the quantity of coal consumed in Cleveland included 117,157 tons of anthracite and 924,602 tons of bituminous coal. Assuming one per cent. of sulphur in the latter, which is the average composition, 9,246 tons of sulphur

were evolved in the combustion of the soft coal, equivalent to 28,305 tons of sulphuric acid, or nearly eighty tons a day. Without doubt this acid is ultimately neutralized in the atmosphere with the formation of ammonium or sodium salts, but directly it is sent forth it must exist chiefly as the free acid. In the evolution of considerable quantities of sulphur no doubt a portion remains for some time in the atmosphere in the form of sulphurous acid. In the city of Lille, France, as much as two cubic centimeters of sulphurous acid per liter has been found in the atmosphere and the rain water collected contained 0.022 gram per liter.

The ammonia evolved in the combustion of coal is not harmful. Indeed in combination with acids it is a normal constituent of the atmosphere. As one of the ultimate products of the decay of vegetable and animal bodies it is continually set free in this manner and it forms an essential element in the food of plants. The cycle of changes includes the evolution of nitrogen as ammonia from decaying organic bodies and the reabsorption as food by plants from the soils which receive it with the rain. Ordinarily natural conditions are sufficient to complete the transformation of the organic compounds during decay into ammonia in an equal distribution such as exists in the economy of nature. But when accumulation of animal or vegetable organic matter occurs at certain points, as in cities, the oxidizing forces of the atmosphere may not be sufficient to accomplish the vast amount of labor involved in decomposition and purification. On account of the extreme instability of the nitrogenous organic compounds these decompositions take place with great readiness in a limited supply of air but with the evolution of products that are a dangerous pollution in the atmosphere. The evidence of such changes in cities is not far to seek, especially if sanitary regulations are not rigidly enforced. Every stable, imperfect sewer connection, cesspool, sink drain, or offal receptacle readily contributes its share to the atmospheric pollution.

In chemical analysis the nitrogen in such products of decay is evolved as ammonia, and from its source in albumen and allied bodies it is known as albuminoid ammonia. Its presence in water or in the atmosphere may be accurately determined and it

may indicate a dangerous condition since its formation is accompanied with, or perhaps is due to the agency of, organisms which are the means of germinating or spreading disease. As might be expected albuminoid ammonia is usually found in calm weather in the atmosphere of cities.

Ammonia has been mentioned as an ultimate product of decay of nitrogenous organic nature; in the ordinary forms of decay and combustion, ammonia is the chief product, but it is usually attended in combustion with the formation, in small quantities, of nitrous and nitric acids, which may be considered as oxidation products. The extensive knowledge recently acquired, concerning the functions of organized germs, has shown that most forms of decay, and of purification as well, are dependent directly or indirectly upon their influence. Through the agency of the germ known as the niter ferment, for example, ammonia may be converted into nitric acid, and this change is constantly in progress where animal refuse accumulates, as in the vicinity of stables. The magnitude of these changes in nature will be appreciated when it is remembered that most of the niter of commerce has been formed in this manner, either naturally, as in the niter beds in hot countries, or artificially, as in the niter plantations of Sweden. Although nitric acid is the ultimate product of such changes, as an intermediary product, nitrous acid is formed in combustion, and both these acids may be detected in the atmosphere where it is receiving, in large quantities, the products of combustion.

In certain cities, including Cleveland, whose rapid development in wealth and population has been dependent, to a large extent, upon their proximity to immense deposits of cheap fuel, the sooty condition of the atmosphere has rapidly increased within recent years until it has attracted much attention with the feeling that such a condition is not conducive either to personal comfort or to health.

Since, so far as we are informed, a thorough examination of the atmosphere in these cities has not been made, it seemed to us of sufficient interest to undertake a study of the atmosphere of Cleveland with especial reference to the constituents which have been described above, with determinations, so far as practicable, of the quantities present.

As a means of ascertaining the quantity of soot deposited from the atmosphere, and also the amount of sulphuric acid removed from the atmosphere by snow, quantities of snow were collected for examination at the following localities which were assumed to represent an average condition of the city atmosphere,—McHenry St., Erie St., Covington St., and campus of the Case School of Applied Science; and the results, with certain data concerning the collection, are given in the following table:

	Location.	Date.	Area.	Weight H_2O from snow melted. Grams.	Weight soot. Gram.
(1)	McHenry St.,	Feb. 19,	9 sq. by 1 in. dp.,	3583	0.3117
(2)	Erie St.,	Feb. 16,	Col. on roof,	3775	0.1709
(3)	Covington St., Newburg,	Feb. 19,		3717	0.4143
(4)	C. S. A. S.,	Mar. 22,	$6\frac{1}{4}$ sq. ft. by 1 in. dp.,	3850	0.1611
(5)	" "	Mar. 27,	$7\frac{1}{2}$ sq. ft. by $\frac{3}{4}$ in. dp.,	2000	.....

	Weight soot per liter.	Weight H_2SO_4 per liter.	Weight NH_3 per liter.
(1)	0.0875 gram	0.0152 gram	0.00007 gram
(2)	0.0452 "	0.0063 "	0.00001 "
(3)	0.1113 "	0.0212 "	0.00012 "
(4)	0.0418 "	0.0139 "	0.000003 "
(5)	.....	0.0070 "	.....

These numbers are interesting in showing the large quantity of soot that falls from the atmosphere with the larger amounts collecting at certain points than at others. The quantities of sulphuric acid are large at all points, but nearly three times greater at Covington St., Newburg, near large establishments, than in the more open area near the Case School of Applied Science. The determinations of ammonia indicate an excessive amount at Covington St., and the least at the Case School of Applied Science.

The specific gravity of the water obtained by melting the snow was also determined:

(1)..... 1.0010
(2)..... 1.0000
(3)..... 1.0010
(4)..... 1.0005
(5)..... 1.0000

These specimens were collected after recent falls before the snow had been contaminated to any considerable extent. The quantity collected on the roof at Erie St. had fallen only a few hours previous to its collection.

On account of the absorptive property of soot it seemed probable that it should retain a considerable quantity of sulphuric acid after it had fallen from the air. For definite proof a quantity of soot was collected on the window sills in the attic of the chemical laboratory, where it had been blown in under the window from the outside. It had probably not been long in collecting. When moistened with water the soot gave a strong acid reaction on test paper, and analysis showed the presence of sulphuric acid.

1.3355 grams of the soot gave 0.00698 gram barium sulphate, equivalent to 1.91 per cent. sulphuric acid.

Another determination of sulphuric acid was made in soot collected in the chimney of a house.

0.4381 gram of this soot gave 0.0459 gram sulphuric acid, which corresponds to 3.17 per cent. sulphuric acid.

Sulphuric acid has invariably been found in testing soot collected in other places.

A noticeable characteristic of a sooty atmosphere and perhaps one of the most important, is the large quantity of soot deposited on the trunks and leaves of trees. In the absence of rain, soot must collect upon the leaves in sufficient quantity to interfere with their natural function of absorbing carbon dioxide. Since it appears that soot and dust absorb sulphuric acid, to a certain extent, from the atmosphere it seemed desirable to ascertain to what extent, if any, this acid collects on the bark of trees. At a point not far removed from manufacturing establishments quantities of bark were scraped from the trunk of a peach, a poplar, a cherry, a hickory, and an oak tree. Probably the kind of tree should not affect the deposition and absorption of acid except perhaps in the degree of roughness of the bark. Nevertheless there are decided differences in the amount of acid from the various sources. The following table gives the results of these determinations :

Variety of trees.	Area.	Total Wt. of bark.	Weight taken for analysis.	Wt. $BaSO_4$.	Wt. H_2SO_4.	Wt. H_2SO_4 per sq. foot.
Peach	245.3 sq. in.	17.5750	5.0008	0.0597	0.0251	0.0148
Poplar	280.0 "	3.7534	3.7534	0.0658	0.0277	0.0132
Cherry	200.0 "	2.8886	2.9896	0.0732	0.0308	0.0221
Hickory	400.0 "	3.1333	3.1333	0.0467	0.0196	0.0071
Oak	400.0 "	6.7886	5.0297	0.0352	0.0148	0.0053

In the order of the roughness of the bark the trees examined may be arranged as follows, with the peach as the least rough: peach, poplar, oak, hickory, cherry. Doubtless the deposition of soot containing acid upon the leaves is of even more serious consequence than upon the bark and the effects are clearly apparent on trees situated in the worst portions of a sooty atmosphere. The foliage of such trees exhibits an appearance of decay late in the summer months and the leaves fall early. We did not succeed in collecting specimens of leaves under suitable conditions to determine to what extent acid is retained on their surface.

Some attention has been given to the condition of the trees in such an atmosphere, but more especially with reference to the ravages of parasites that infest trees. In individual instances, no doubt, trees may be totally destroyed in this manner, but it is questionable whether the pernicious effects of a sooty atmosphere are not of more serious consequence on the trees as a whole. The corrosive action of an atmosphere impregnated with soot is daily manifest to the most ordinary observation. I have seen, in the country, nails withdrawn from a shingled roof seventy-five years after the shingles were laid, with the heads apparently unaffected, and the portion between the shingle and the board only slightly rusted away. In the atmosphere of a sooty city, as every one knows, after ten or fifteen years the nails on a roof are nearly consumed.

The effects of such an atmosphere on perishable property, such as certain forms of merchandise, are not so readily discernible, since this material is soon consumed by use. It is in libraries that the effects are most apparent. Fifteen years ago it was demonstrated by the late Professor W. R. Nichols that sulphuric acid is absorbed in bookbindings, and after some time the bindings begin to crumble; he found that the presence of the acid in deteriorated bindings may readily be shown.

The samples of air for examination, were collected at the following points, as fairly representing the different sections of the city.

(1) Campus of Case School of Applied Science, not far from the school buildings and those of Adelbert College. The line of the N. Y., St. Louis & Chicago R. R. is one-eighth of a mile distant toward the south-east.

(2) McHenry St.; closely inhabited, numerous manufactories 700–1,000 feet distant; a line of foundries, steel works, etc., extends three miles along the lake front, one-half a mile distant toward the north from this point. Most of the solid particles of soot fall to the ground within one-fourth to one-half a mile from the source.

(3) Rockwell St.,; surrounded by large business blocks. The general condition of the atmosphere is bad on account of the soot, dust from the surrounding streets, and laxity in the enforcement of sanitary regulations.

(4) St. Clair St., near St. Lawrence; many large manufacturing establishments within 500 feet, on the lake shore, all using bituminous coal. When all are in operation, dense volumes of sooty smoke are sent forth into the atmosphere.

(5) West side; on the brow of the hill above the river near Pearl and Detroit Sts. Many locomotives pass below during the hour, and, on the river, tugs pass frequently. When the wind blows from the south-east, this locality is probably one of the worst in the city. The river itself doubtless aids the atmospheric pollution at this point since it is little better than an open sewer with a very sluggish current.

(6) South side, Jefferson St., a residence section with no factories near. In a north-easterly direction is a large ravine containing various kinds of decaying organic matter.

(7) Valley of the Cuyahoga, "Flats"; a very large number of factories of all kinds together with slaughter houses, rendering establishments, etc. Noxious vapors and disagreeable odors from the river invest this entire locality. In clear weather the smoke ascends for the particular benefit of the inhabited districts on the neighboring bluffs. Circulation of the atmosphere is

slow in damp weather and the impure atmosphere remains in the valley.

(8) Genessee Avenue, near Superior St.; a residence and farming district with no factories nearer than one-fourth of a mile.

(9) Mayfield, three miles outside of the city limits and six miles from the large manufacturing establishments of the city; a farming section with no barns nearer than one-fourth of a mile of the point where the collection of air was made.

(10) Collinwood, three miles from the city limits, near the Lake Shore Railroad, where locomotives are constantly passing; numerous barns close together, the nearest 900 feet distant.

Numerous determinations of carbon dioxide at different points in sections of country widely separated have shown no appreciable variation. On the average it is not far from four parts in 10,000 with a variation between three and five-tenths and four and five-tenths parts. Analyses of Cleveland air have shown some variation evidently depending upon the particular conditions at the point where the collection was made:

Locality.	Parts CO_2 per 10,000.
(1)	3.52
(2)	3.89
(3)	4.01
(4)	3.86
(5)	4.52
(6)	4.31
(7)	4.77
(8)	3.99
(9)	3.68

We have attached no particular importance to the determination of carbon dioxide in connection with the principal object of this examination. The following numbers represent the quantity usually present in a city atmosphere:

	Parts in 10,000.
London, England	4.39
Glasgow, Scotland	5.02
Perth, "	4.12
Manchester, England	4.42
Cleveland	4.42

In determining soot in the atmosphere a measured volume of

air was drawn through a piece of combustion tubing partly filled with ignited asbestos and the carbon was determined by ignition, and absorption of the carbon dioxide as in ordinary combustions. The weights given in the following table are probably somewhat too low, since the soot is doubtless composed, for the most part, of hydrocarbons:

Locality.	Weather.	Wind	Vol. of air at 0° and 76 cm. Liters.	Wt. CO_2.	Weight soot in 1,000,000 cubic meters.
(1)	Clear	N.	94.0	0.0034 gram	9,875 grams
(2)	"	S. W.	70.2	0.0017 "	6,609 "
(3)	"	N. W.	88.2	0.0051 "	15,773 "
(5)	"	S. W.	92.5	0.0042 "	12,386 "
(5)	"	N. W.	88.7	0.0005 "	1,538 "
(6)	Cloudy	W.	89.3	0.0054 "	16,491 "
(6)	Clear	W.	91.3	0.0027 "	8,065 "
(8)	"	N. W.	94.3	0.0053 "	15,330 "
(8)	"	N. W.	77.9	0.0140 "	39,906 "

A clearer idea may be gained of the quantities of soot given above, if it is understood that 1,000,000 cubic meters is equivalent to the quantity of air contained within a space one mile square and a little more than one and one-fourth feet deep, and that the quantity of carbon, 39.664 grams calculated from analysis of the air collected at Genessee Ave., is equivalent to approximately 8.7 pounds. This very high result has no direct bearing on the quantity of soot in the atmosphere at this point, since it may be attributed to a high wind which filled the air with dust while the experiment was in progress. The air was collected at McHenry St., seven hours after rain had fallen. In certain conditions of the atmosphere frequently observed, there can be no doubt that the quantity of organic dust in suspension is very much larger than is represented by the values given above. The variable proportion of carbonaceous matter in the city air is shown in the duplicate samples collected at the West Side, the South Side, and Genessee Avenue. With the wind blowing from the north-west at the West Side the air is very clear, but in the direction of the south-east, as mentioned above, this section is in the direct line of the smoke from the river valley. Evidently these determinations serve to indicate merely the quantity in suspension in the air examined; but since the soot is rapidly

deposited, these values are, in no sense, a measure of the quantity escaping into the atmosphere.

In determining sulphur in the atmosphere in the form of sulphuric acid, a measured volume of air was drawn through a large U tube containing a solution of sodium hydroxide and broken glass, after passing through ignited asbestos. The alkaline solution was acidified with hydrochloric acid, oxidized with bromine, and the sulphuric acid precipitated with barium chloride. The quantities obtained are given in the following table:

	Date.	Locality.	Weather.	Wind.	Velocity of air.
(1)	Apr. 9, A. M.	C. S. A. S.	Clear	N. E.	53.9 liters
(2)	" 9, P. M.	"	"	Variable	53.9 "
(3)	" 13,	"	"	N. E.	6.25 miles 101.5 "
(4)	" 19,	McHenry St.	Cloudy	S. W.	3.7 " 27.3 "
(5)	June 13,	" "	"	W.	1.5 " 55.7 "
(6)	Apr. 23,	Rockwell "	"	N. W.	29.4 "
(7)	" 24,	St. Clair "	Hazy	N. W.	3.6 " 28.8 "
(8)	" 27,	Riv. Val.	Clear	S. E.	5 " 27.6 "
(9)	June 29,	" "	Cloudy	N.–S.	5.5 " 47.8 "
(10)	Apr. 30,	W. Side	Clear	S. W.	1.5 " 28.3 "
(11)	May 1,	S. Side	Cloudy	W.	4.5–18.3 " 28.0 "
(12)	" 7,	Genessee Ave.	"	N. W.	8.7–2.5 " 28.6 "
(13)	" 15,	Collinwood	Hazy	N. E.	44.2 "
(14)	June 26,	Mayfield	Clear	S. W.	7.5 " 39.8 "

Locality.	Weight $BaSO_4$.	Grams H_2SO_4 in 1,000,000 cu. met.
(1)	0.0030	23414
(2)	0.0028	21859
(3)	0.0053	20178
(4)	0.0027	42156
(5)	0.0022	16615
(6)	0.0033	47801
(7)	0.0011	16142
(8)	0.0033	50629
(9)	0.0033	29014
(10)	0.0057	56828
(11)	0.0042	63289
(12)	0.0023	33989
(13)	0.0000	.....
(14)	0.0007	7392

The variation in the quantity of sulphuric acid in the atmosphere is shown in the duplicate determinations at McHenry St., and the river valley, made at different dates. A negative result

at Collinwood is, perhaps, what might be expected in country air under certain conditions. And yet it seems that it was possible for sulphuric acid to be carried as far as Mayfield. It would be interesting to examine the atmosphere at different distances outside the city to ascertain how far the influence of the city air extends.

The following weights in grams of sulphuric acid in 1,000,000 cubic meters, have been found in the atmosphere of cities in England:

London	1669.935
Didsbury	1337.844
Blackpool	474.413
Buxton	1494.400
Manchester	2443.738

As already explained, sulphur sent forth into the atmosphere from the combustion of coal may escape complete oxidation and appear in part as hydrogen sulphide and in part as sulphurous acid. In testing for sulphurous acid in the atmosphere near the Case School of Applied Science, a measured volume of air was drawn through the alkaline solution for absorption, and in one portion the sulphuric acid was precipitated with barium chloride. The weight of barium sulphate corresponded to 20,276.36 grams of sulphuric acid in 1,000,000 cubic meters. To the remaining portion of the solution used for absorption bromine was added for oxidation, and the solution precipitated with barium chloride. The difference in weights corresponded to 33,033.1 grams of sulphuric acid in 1,000,000 cubic meters. In further confirmation of the presence of both sulphurous and sulphuric acids, another determination was made on the "Flats," near the Central viaduct, on an exceptionally clear day. In eighty-six liters of air aspirated, the total weight of barium sulphate, after oxidation, corresponded to 0.0058 gram barium sulphate, and the weight of barium sulphate before oxidation was equivalent to 0.0028 gram barium sulphate, the difference, 0.0031, representing the sulphurous acid in the total volume of air aspirated. The quantity of sulphuric acid present was, therefore, represented by 11,851 grams, and the sulphurous acid by 12,512.6 grams in 1,000,000 cubic meters.

In the atmosphere of other cities, notably in Manchester,

England, and Lille, France, sulphurous acid has been determined, and it is probable that it may always be detected in the atmosphere where it receives large quantities of smoke from the combustion of coal.

The presence of ammonia as a normal constituent of the atmosphere has already been mentioned, and under ordinary conditions it does not collect in sufficient quantity to be injurious. Nevertheless, it seemed of interest to ascertain to what extent this constituent collects in an atmosphere contaminated with soot. In collecting ammonia, a definite volume of air was drawn through a dilute solution of hydrochloric acid, and both free and albuminoid ammonia were determined by the methods ordinarily employed in water analysis. After the free ammonia was distilled the residual solution was oxidized with alkaline permanganate until there was no further evolution of ammonia, and both sets of distillates nesslerized. The following tables give the quantity of free and albuminoid ammonia from the quantity of air aspirated in each locality and the state of the weather.

	Date.	Locality.	Weather.	Wind Direc.	Wind Veloc.	Volume of air in liters.
(1)	Apr. 6,	C. S. A. S.	Clear	N. W.	9 m.	48.024
(2)	" 6,	"	Cloudy	"	6 "	46.071
(3)	" 19,	McHenry St.	"	S. W.	4 "	20.391
(4)	June 13,	" "	"	W.	1.5 "	6.941
(5)	Apr. 23,	Rockwell "	"	N. W.	4 "	22.334
(6)	" 24,	St. Clair "	Hazy	N. W.–W.	2 "	22.163
(7)	June 20,	" " "	Clear	N.		6.963
(8)	Apr. 30,	West Side	"	S. W.	18 "	21.282
(9)	" 27,	" "	"	N. W.	1 "	7.054
(10)	May 1,	South Side	Cloudy	W.	1 "	20.971
(11)	June 22,	" "	Clear	W.	0.5 "	7.092
(12)	May 7,	Genessee Ave.	Cloudy	N. W.–S. W.	8.7–25m	6.622
(13)	June 21,	" "	Clear	N. W.	1 m.	6.829
(14)	" 26,	Mayfield	"	S.	7.5 "	6.781
(15)	" 29,	River Valley	"	N. W.	6.5 "	6.862

	Total free NH_3.	Total Albuminoid NH_3.	Free NH_3 in 1,000,000 cu. met.	Albuminoid NH_3 in 1,000,000 cu. met.
(1)	0.00000051	0.0000071	106.10 grams	147.84 grams
(2)	0.00000490	0.0000045	106.35 "	97.67 "
(3)	0.00002180	0.0000113	1070.20 "	557.50 "
(4)	0.00000210	0.0000021	302.50 "	298.90 "
(5)	0.00000780	0.0000035	357.20 "	158.90 "
(6)	0.00002330	0.0000051	1051.10 "	230.10 "

	Total free NH_3.	Total Albuminoid NH_3.	Free NH_3 in 1,000,000 cu. met.	Albuminoid NH_3 in 1,000,000 cu. met.
(7)	0.0000014	0.0000023	198.7 "	329.1 grams.
(8)	0.0000082	0.0000037	385.5 "	173.8 "
(9)	0.0000059	0.0000030	836.9 "	430.9 "
(10)	0.0000169	0.0000043	806.4 "	202.5 "
(11)	0.0000015	0.0000007	214.5 "	100.6 "
(12)	0.0000086	0.0000031	1373.0 "	489.5 "
(13)	0.0000030	0.0000034	440.6 "	492.0 "
(14)	0.0000035	0.0000019	510.2 "	282.8 "
(15)	0.0000054	0.0000010	787.7 "	150.1 "

We have found on record but few determinations of free and albuminoid ammonia, but similar determinations have been made in certain cities abroad, the results of which afford an interesting comparison with those we have obtained.

	Grams free NH_3 in 1,000,000 cu. met.	Grams albuminoid NH_3 in 1,000,000 cu. met.
London	61.286	150.919
Glasgow	78.196	304.974
Innellan	52.281	137.832

These numbers represent much smaller quantities of albuminoid ammonia, and still smaller quantities of free ammonia than appear in the results of our determinations.

It has long been known that nitrous acid and nitric acid are contained in the products of combustion from fuels containing nitrogen as most fuels do. These acids represent a more advanced stage of oxidation than ammonia ; they are derived to a certain extent, it may be, from the oxidation of atmospheric nitrogen ; and they may also be detected in the products of respiration. Ammonium nitrate may be found in dew and it seems to be present in larger quantities in spring than in summer and more at night than during the day. These acids are, therefore, normal constituents of the atmosphere under natural conditions, and the minute quantities usually present can not be considered harmful.

In testing for nitric acid we used as an indicator phenolsulphonic acid prepared by dissolving phenol in concentrated sulphuric acid and for comparison, a solution of potassium nitrate in which one cc. contained 0.000001 gram nitrogen. For nitrous acid, sulphanilic acid and naphthylamine hydrochloride were used as indicators with a solution of sodium nitrite, in which one cc. contained 0.000001 gram nitrogen as a means of com-

parison. A measured volume of air was drawn through a solution of sodium hydroxide prepared from metallic sodium, and definite quantities of this solution were used in each determination.

The following tables give the quantities of nitrous and nitric acids found in experiments conducted at different localities with other data of the collection:

NITROUS ACID.

Date.	Place.	Weather, &c.		Grams HNO_2 in 1,000,000 cub. met.
Apr. 17,	C. S. A. S.	Clear	W. N. E.	6.767
May 11,	"	"	W. N. W.	0.000
July 2,	"	Cloudy	Calm.	6.955
April 19,	McHenry St.	"	W. S. W.	119.277
June 30,	" "	Clear	W. N. W.	73.976
April 23,	Rockwell "	Cloudy	Damp	152.785
June 16,	" "	Clear	W. N. W.	8.856
Apr. 24,	St. Clair "	"	W. N. W.	126.614
June 20,	" "	"	W. N.	135.500
June 29,	River Valley	Cloudy	W. N. W.	0.607
May 4,	West Side	Clear	W. N. W.	15.206
June 27,	"	"	W. N. W.	19.945
June 22,	South Side	"	W. W.	1.239
May 7,	Genessee Ave.	"	W. N. W.	5.833
June 21,	" "	"	W. N. W.	7.228
May 15,	Collinwood	Cloudy	Calm.	4.564

NITRIC ACID.

Date.	Place.	Grams HNO_3 per 1,000,000 cub. met.
April 17,	C. S. A. S.	674.5
May 11,	"	560.0
April 19,	McHenry St	703.8
June 13,	" "	54.99
April 24,	St. Clair St	520.5
June 20,	" "	144.0
April 27,	River Valley	704.0
June 27,	" "	68 8
April 30,	West Side	1063.6
May 4,	" "	367.6
June 27,	" "	66.9
May 1,	South Side	61.8
June 22,	" "	153.8
May 7,	Genessee Avenue	333.6
June 21,	" "	129.4
May 15,	Collinwood	424.0
June 16,	Rockwell St	544.8
June 26,	Mayfield	44.5

The variation in the proportion of both nitrous and nitric acids under different conditions is very marked, as shown in the samples of air collected in the river valley, the West Side, and McHenry St.

From the results of this examination it is evident that a city atmosphere, contaminated by the universal consumption of bituminous coal, where no efforts are made to prevent the escape of soot, soon reaches a stage in which it is destructive to property and not conducive to health. In this respect the atmosphere of Cleveland is, doubtless, no worse than that of other cities, and, perhaps, in a better condition than some that use the same fuel. Under the usual conditions of life in cities, sanitary regulations require careful attention and constant supervision.

A NEW FORM OF WATER-OVEN AND STILL.[1]

By Lewis William Hoffmann and Robert W. Hochstetter.

Received November 8, 1894.

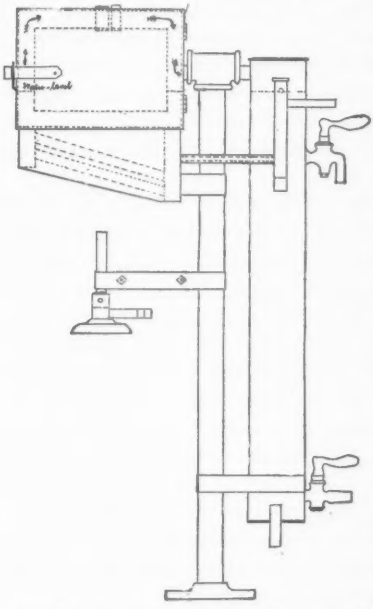
In the use of the ordinary form of water-oven we have been impressed with the fact that there are several disadvantages connected with that form of this apparatus. For instance, there is the necessity for refilling it from time to time, with the consequent fall in temperature and loss of time; there is the waste of steam which might be condensed to the always-useful distilled water, and we have had trouble in obtaining a constant temperature of 100° C. To obviate these difficulties we have devised the combined oven and still which we describe here—a piece of apparatus that has an automatic feed, a worm for condensing the steam and an arrangement of the heating surface so as to provide for perfect circulation and quiet boiling, with steady and abundant generation of steam.

The combination of still and oven is not new, for in the April, *1892*, number of the *Journal of Analytical and Applied Chemistry*, Herbert M. Hill describes such a piece of apparatus in use in the Buffalo University laboratory. Except for the combination it bears little resemblance to our apparatus.

In our form of apparatus, the heating surface consists largely of a series of pipes, connected with water boxes at each end.

[1] Read before the Cincinnati Section, October 15, 1894.

These pipes, through which the water circulates, are inclined so as to afford an easy ascent for the steam generated in them, as the result of the heat applied to their outer surface. The source of heat is an ordinary Bunsen burner, used with a wing-tip, so as to spread the flame over the surface of the pipes. The pipes are of small diameter so as to afford a large heating surface in proportion to the amount of water passing through them.



The steam is carried off at a point about half-way up the side of the oven and passed through a block-tin condenser. To reach the condenser the steam must pass over the top of the drying-chamber, as the water-level in the oven is kept about one-fourth of the distance between the top and bottom, thus acting as a lute and compelling the steam to pass up and over in order to escape. The water used in cooling the worm also serves as feed-water for the oven, and as it passes from the bottom to the top of the condenser, absorbs so much heat as to enter the water-box already heated almost to the boiling-point. The feed-pipe is so small that the circulation between condenser and oven is slow but constant. A stop-cock is provided at the top of the condenser, for drawing off hot water if desired. A small hole in the door and another in the top provide for ventilation of the drying chamber. Sedimentation takes place in the condenser so that almost clear water goes into the heating-pipes.

To the factory chemist, having always a plentiful supply of condensed steam at his disposal, this piece of apparatus may not seem of much value, but to the analyst, a combination of

still and water-oven means a saving of time, trouble, and, in some cases, expense.

The advantage derived from the arrangement of pipes for receiving heat, (an idea borrowed from the Babcock-Wilcox boiler,) may be conceived from the statement that while in the oven which we have, the heating surface measures 82.8 square inches, but seven square inches of this area is the bottom of the oven, the rest being made up by the pipes and the sides of the water-boxes at each end.

Now as to the work done by the apparatus. The oven which we have in our laboratory will get up steam in four minutes after the flame is lighted, and at the expiration of six and two-tenths minutes will be delivering distilled water. This is with a Bunsen burner consuming three cubic feet of gas per hour. With a burner consuming twice as much gas, one and three-tenths minutes is gained in steaming, and two and two-tenths minutes in delivering water.

The still delivers water at the rate of two and one-fourth quarts per hour or thirteen and one-half gallons in twenty-four hours—this with a three-feet gas-burner. A patent automatic still which we have hitherto been using, consumes six cubic feet of gas per hour, and delivers two quarts of water an hour, or twelve gallons in twenty-four hours, although the claim made for it by the makers is fifteen gallons in twenty-four hours.

The temperature in the drying chamber of our form of apparatus ranges from 98° to 101° C., so that it may be said to average 100°.

The advantages claimed for our form of apparatus are as follows:

A constant temperature of 100°, which in our experience, can not be obtained with the ordinary form of oven.

A large yield of distilled water, with a small expenditure of heating material, and practically no cost, as the water-oven is nearly always in use in a working laboratory.

Hot water always at command.

Small space occupied by both oven and still.

General convenience of arrangement.

MT. AUBURN, CINCINNATI,
November, 1894.

[CONTRIBUTIONS FROM THE CHEMICAL LABORATORY OF THE UNIVERSITY OF MICHIGAN.]

THE DETERMINATION OF NICKEL IN NICKEL-STEEL.

E. D. CAMPBELL AND W. H. ANDREWS.

Received December 24, 1894.

SINCE publishing his former method for the determination of nickel in nickel-steel (This JOURNAL, **16**, 96), the author has been endeavoring to perfect another method which should be an improvement over the first one. Such accurate results have been obtained in so short a time by this method that it is thought desirable to give the details of manipulation.

Place a one gram sample of nickel-steel in an Erlenmeyer flask of about one-half liter capacity. Add twenty-five cc. of nitric acid (1.20) and boil until dissolved and nitric oxide is boiled out of the flask. Meanwhile, dissolve thirteen grams of sodium pyrophosphate in seventy-five cc. of warm water. This solution must not be boiled as the pyrophosphate is thereby converted to the orthophosphate.

After complete solution, filter the sodium pyrophosphate and add it to the cooled solution of the steel. This will give a heavy, white precipitate of ferric pyrophosphate. Add, a little at a time, with constant agitation, a clear twenty per cent. solution of sodium carbonate until the precipitate of ferric pyrophosphate just dissolves and the solution attains a dull, olive-green color.

An excess of sodium carbonate must be carefully avoided, as the subsequent precipitation of nickel will not be complete if this is not done. With an excess of the carbonate the solution turns red, so that color may be used as a guide to the required amount. For a one gram sample, fifty to fifty-five cc. of twenty per cent. sodium carbonate is usually sufficient.

After redissolving with sodium carbonate, filter through asbestos, receiving the solution in a one-half liter Erlenmeyer flask; wash the filter with water sufficient to rinse out the original flask and to insure complete transference of the solution. Now add to the solution two grams of potassium xanthate freshly dissolved in twenty to twenty-five cc. of water. Stopper the flask and shake at frequent intervals for ten minutes. The nickel will be

completely precipitated as xanthate, together with any copper the steel may contain, but perfectly free from manganese and almost free from iron. Filter as rapidly as possible through asbestos, using a pump, and wash the brick-red nickelous xanthate four or five times with freshly prepared one per cent. potassium xanthate.

The solution of potassium xanthate used for precipitating, as well as that used for washing, should be prepared by adding to the desired amount of salt sufficient water for solution, but should not be allowed to stand longer than fifteen or twenty minutes before using as it slowly decomposes. The dry salt is stable.

Dissolve the washed precipitate off the filter by allowing twenty cc. of freshly diluted fuming nitric acid to fall little at a time on it. The dilute fuming nitric acid should be prepared just before using by pouring ten cc. of water into ten cc. of fuming nitric acid. Catch the solution of the xanthate precipitate in a small Erlenmeyer flask. Wash the filter two or three times with hot water, then pass through five cc. of dilute sulphuric acid (2:3) and wash this out with more hot water. Boil the solution in the flask rapidly until the nitric acid is completely expelled and sulphuric anhydride fumes begin to rise. Cool, add twenty to twenty-five cc. of water, heat nearly to boiling, and, when all is in solution, pass in hydrogen sulphide for two or three minutes. Allow to stand until the cupric sulphide has settled. Filter and wash with hot water containing a little hydrogen sulphide.

Boil the filtrate, containing the nickel and a very small amount of iron, until hydrogen sulphide is expelled, add a few drops of hydrogen peroxide, or a few milligrams of sodium peroxide to oxidize the iron, and boil, then precipitate the iron with a very slight excess of ammonium hydroxide. Boil, filter, and wash once or twice with hot water. Dissolve the small precipitate of ferric hydroxide off the filter with five cc. of dilute sulphuric acid (2:3) and reprecipitate with a slight excess of ammonium hydroxide. Boil and filter into the original filtrate. This solution will contain all the nickel free from iron, manganese, and copper The nickel may be determined in either of the following ways :

1. *Electrolytically.*—Place the united filtrates from the ferric hydroxide in a platinum dish having a capacity of about 200 cc., add thirty cc. of a ten per cent. solution of disodium hydrogen phosphate, then twenty-five cc. of strong ammonium hydroxide. The solution should have a bulk of about 175 cc. Electrolyze for twelve hours with a current of 0.14 amperes per 100 sq. cm. area of the dish, having the electrodes about one-half cm. apart.

The precipitate of nickel obtained in this way is perfectly clear and bright, hardly to be distinguished from the platinum on which it is deposited.

Wash the precipitated nickel with warm water then with alcohol. Dry and weigh as usual. For accurate work it is more satisfactory to use a second platinum dish washed and dried like the first, as a counterpoise.

2. *Volumetrically.*—Place the united filtrates from the ferric hydroxide in a medium sized Erlenmeyer flask and cool until cold. The solution should have just enough free ammonia to give a slight but distinct odor; if this is not present at this stage, add a few drops of ammonium hydroxide. Add five cc. silver nitrate indicator (made by dissolving five-tenths gram silver nitrate in 1,000 cc. of water). Add five cc. of a two per cent. solution of potassium iodide, which give the solution an opalescence due to the presence of silver iodide in suspension.

Run in standard potassium cyanide (one cc. = 0.001 gram nickel) until the solution becomes clear and bright. The titration depends upon the fact that so long as there is any nickel uncombined with potassium cyanide in the form of the double cyanide, the silver iodide will remain undissolved, but as soon as the nickel has all been changed the silver iodide dissolves forming potassium silver cyanide and the solution becomes perfectly clear.

Subtracting a correction necessary for the titration of the indicator from the total amount of potassium cyanide used, gives the amount necessary for the nickel.

Standard Solutions Required.—Standard nickel solution: This can be most easily prepared from pure nickel wire. The commercially pure nickel wire usually contains 98.5–99.5 per cent.

nickel. If this wire is used for making the standard solution, take such a weight as will represent one gram of pure nickel, dissolve in a No. 3 beaker with twenty-five cc. nitric acid (1.42) and evaporate just to dryness on the water-bath. Moisten with a few drops of nitric acid, add about twenty-five cc. of water, and heat to solution. When the nickel is dissolved, dilute to fifty or sixty cc., add six grams of sodium acetate dissolved in a little water, heat nearly to boiling until the ferric basic acetate separates. Filter, and wash thoroughly with hot water. Add five cc. of sulphuric acid to the solution, transfer to a graduated liter flask, cool, and dilute to 1,000 cc.

Standard Potassium Cyanide.—Dissolve four and five-tenths grams of potassium cyanide in water and dilute to 1,000 cc. This must be standardized against the standard nickel solution.

Run in from a burette twenty-five to thirty cc. of the nickel solution into a medium sized Erlenmeyer flask, dilute with water to about 100 cc., add ten cc. of dilute sulphuric acid (2:3); then add ammonium hydroxide carefully until a slight but distinct odor of ammonia is perceptible. Add five cc. of silver nitrate indicator, five cc. of two per cent. potassium iodide and titrate with potassium cyanide until the solution brightens. Correction for the indicator is found by proceeding exactly as above, except that the standard nickel solution is omitted.

The potassium cyanide is quite stable, weakening, however, gradually on standing. The deterioration amounted from five-hundredths to one-tenth cc. in twenty-five cc. during one week.

In order to test fairly the accuracy of the recovery, the varying amounts of pure nickel wire were weighed and dissolved by my private assistant, who turned them over to Mr. Andrews as strictly unknown samples. Weighed amounts of iron wire were then added and the solution proceeded with as if it were an ordinary sample of steel.

Two hours is sufficient time to get the solution ready for electrolysis or for titration.

The following table will show the accuracy of the results obtained, where the amount of nickel present was strictly unknown to the operator until the analysis was completed:

Iron added.	Nickel taken.	Nickel recovered.	Gain and loss. + −	Method.
Grams.	Gram.	Gram.	Gram.	
0.2105	0.08892	0.08910	+ 0.00018	Electrolytically.
1.0000	0.03430	0.03435	+ 0.00005	"
1.0041	0.05298	0.05345	+ 0.00047	"
1.0043	0.10824	0.10825	+ 0.00001	"
2.0043	0.02690	0.02740	+ 0.00050	"
1.0010	0.00100	0.00089	− 0.00011	Volumetrically.
1.0086	0.06578	0.06532	− 0.00046	"
1.0066	0.07582	0.07583	+ 0.00001	"
1.0057	0.08017	0.08050	+ 0.00033	"
1.0071	0.09389	0.09381	− 0.00008	"

ANN ARBOR, MICHIGAN,
December 18, 1894.

[CONTRIBUTIONS FROM THE CHEMICAL LABORATORY OF THE ROSE POLYTECHNIC INSTITUTE. VI.]

THE VOLUMETRIC DETERMINATION OF PHOSPHORUS IN STEEL AND CAST IRON.[1]

BY W. A. NOYES AND J. S. ROYSE.

Received January 3, 1895.

IN a previous paper[2] one of us, with Mr. E. D. Frohman, showed that in the reduction of molybdic solutions by means of Jones' reductor, the molybdenum is reduced to a form corresponding to the oxide Mo_2O_3. We also showed that the reduced solution is very easily oxidized by exposure to the air, and we were led to believe this would account for the difference between the theoretical factor for the calculation of the phosphorus equivalent of a permanganate solution from its iron equivalent, and the empirical factor as given by Dudley and Pease.[3]

In a private communication, however, Dr. Dudley has very kindly called our attention to the fact that he obtained the same results when following the directions for reduction given in our paper as when following the method which he had previously used. As the difference between his results and ours corresponded to only three-tenths or four-tenths cc. of the permanganate solution used, and the results obtained by Mr. Frohman

[1] Read at the Boston Meeting, December 28, 1894.
[2] This JOURNAL, 16, 553.
[3] This JOURNAL, 16, 230.

were not sufficiently accordant to establish the point in question, it became necessary to repeat that part of the work. Accordingly, the following series of determinations was made, following, as closely as possible, the directions given by Dudley and Pease.[1] The results are given in per cents., on the basis of one gram of steel.[2]

	Phosphorus present.	Phosphorus found. By theoretical factor.	Phosphorus found. By factor of Dudley and Pease.
Steel No. 1	0.040	0.040	0.045
" " 1	0.040	0.0354	0.0397
" " 2	0.053	0.0525	0.059
" " 2	0.053	0.0525	0.059
" " 3	0.032	0.0264	0.0296
5 grams cryst. $FeSO_4$ + 0.0011 gram cryst. Na_2HPO_4	0.010	0.0035	0.0039
5 grams cryst. $FeSO_4$ + 0.0130 gram cryst. Na_2HPO_4	0.113	0.127	.....
5 grams cryst. $FeSO_4$ + 0.0097 gram cryst. Na_2HPO_4	0.084	0.091	.....
5 grams cryst. $FeSO_4$ + 0.0109 gram cryst. Na_2HPO_4	0.094	0.106	.....
1 gram steel No. 3 + 0.0068 gram cryst. Na_2HPO_4	0.091	0.097	.....
1 gram steel No. 3 + 0.0026 gram cryst. Na_2HPO_4	0.055	0.054	0.0061

The first result is evidently wrong for some unknown reason. Omitting this, the results with steels Nos. 1 and 3 show conclusively that the factor of Dudley and Pease is correct for those steels and the statement of our first paper was erroneous. As regards steel No. 2 the case is different. Dudley and Pease explain the higher results obtained with this steel as due to arsenic which it contains. The results of the last determination in the table, and, still more, the results obtained by another method and given below, lead us to believe that this is not the case, but that for a steel containing 0.055 of phosphorus the theoretical factor gives very nearly correct results.

From the examination of the table as a whole it becomes apparent, at once, that, while the factor given by Dudley and Pease gives results nearly or quite correct for steels containing 0.03 to

[1] *J. Anal. Appl. Chem.*, 7, 108. This JOURNAL, 15, 519.

[2] This set of determinations was made by W. A. Noyes.

0.04 per cent. of phosphorus, it fails entirely so soon as the amount present is either above or below these limits. The reason is, we believe, apparent. When the amount of phosphorus is less than 0.05 per cent., a part of it escapes precipitation, and hence, a larger factor than the theoretical is required. When the amount of phosphorus is as great as 0.08 per cent. or more, however, the yellow precipitate probably carries down with it at 55° C., some free molybdic anhydride, and even the theoretical factor gives too high results. If we retain the method in its present form, therefore, it will be necessary to use several different empirical factors, or to apply a correction varying with the amount of phosphorus present in the steel under examination. We believe such a course is open to very serious objection and that it should be avoided, if possible. The use of a variable correction implies that the yellow precipitate is not a pure chemical compound, but that it has a variable composition dependent on the amount of phosphorus present. If we admit this, it is evident that the method is not a suitable one on which to base an accurate determination, for a compound or mixture of variable composition is liable to be seriously affected by even slight changes in manipulation, changes which are sure to occur in different hands and different laboratories.

These considerations led us to endeavor to modify the method in such a manner as to secure, if possible, a practically complete precipitation of very small amounts of phosphorous and also to secure the precipitation in the form of a compound of definite composition and independent of the amount of phosphorus present. The considerations which have guided us in the study are as follows: first, the precipitation would be more rapid in a more concentrated solution; second, the precipitation must be at a lower temperature than 55° to prevent the contamination of the precipitate with molybdic anhydride; third, the precipitate is formed more rapidly in a more nearly neutral solution; and fourth, the presence of ammonium nitrate hastens the formation of the precipitate, probably because ammonium enters into its composition.

Without going into the details of unsuccessful experiments, we will give the method which we have found most satisfactory.

We give details for reagents and apparatus, though most of these details are a repetition of those given by Dudley and Pease.

REAGENTS AND APPARATUS.

Nitric Acid.—One part of concentrated acid (1.42) to two parts of water by volume. This gives an acid of sp. gr. 1.18.

Oxidizing Permanganate Solution.—Twelve and one-half grams of potassium permanganate in one liter of water.

Ferrous Sulphate.—The crystallized non-effloresced salt, *free from phosphorus.*

Ammonia.—Concentrated, sp. gr. 0.90; and dilute, sp. gr. 0.96.

Molybdic Solution.—Dissolve 100 grams of molybdic anhydride in 400 cc. of ammonia (0.96), filter and pour the solution slowly, with constant stirring, into 1,000 cc. of nitric acid (1.20). Warm to 45°, add one cc. of a ten per cent. solution of crystallized sodium phosphate, stir vigorously, and allow to stand in a warm place over night before using. The addition of the sodium phosphate was suggested by Dr. Dudley and has the two-fold object of precipitating anything liable to contaminate the yellow precipitate and of leaving the solution saturated with the compound which is to be precipitated by it. Our experience leads us to think that the solution keeps better in the dark. It is always safer to filter it immediately before it is used and the solution must, of course, be absolutely clear. If pure, it will be almost or quite colorless, but we have found that a somewhat yellow solution prepared as above, gave us results nearly or quite identical with those obtained with a colorless solution.

Acid Ammonium Sulphate.—To one half liter of water add twenty-seven and one-half cc. of ammonia (0.96) and twenty-four cc. of concentrated sulphuric acid and make up to one liter.

Potassium Permanganate for Titration.—Dissolve four grams of potassium permanganate in two liters of water, heat nearly to boiling for an hour, allow to settle over night, and filter with the pump, using an asbestos filter prepared after the manner of Gooch on a small perforated porcelain plate placed in the bottom of a large funnel. Use very great care that the solution does not come in contact with india-rubber or any kind of organic matter and keep in a clean glass-stoppered bottle, free from

manganese oxides. We find that a solution prepared in this manner keeps perfectly and is not affected by the light.

We prefer to standardize with freshly crystallized and thoroughly air-dried ammonium oxalate dissolved in water with the addition of a little dilute sulphuric acid (instead of considerable concentrated acid as advised by Fresenius) and with ferrous ammonium sulphate freshly crystallized in small crystals from a slightly acid solution, the crystals being thoroughly washed and quickly air-dried in a thin layer, using the factors $\frac{112}{142}$ and $\frac{1}{7}$ respectively to calculate the iron equivalent[1]. The phosphorus equivalent is obtained by multiplying the iron equivalent by $\frac{31}{36\times56}=0.01538$.

Jones' Reductor.—A piece of moderately heavy glass tubing thirty-five cm. long and having an internal diameter of about two cm., is drawn out at one end to pass through the stopper of a filtering flask. In the bottom of the tube is placed a small perforated porcelain plate and on this is prepared an asbestos filter, practically a Gooch filter. This filter should offer enough resistance so that a reduction will occupy two or three minutes. The tube is nearly filled with powdered zinc, which is then washed with dilute sulphuric acid (1:20 by volume). Blank determinations are made as follows: To 180 cc. of water, add twelve cc. of ammonia (0.96) and ten cc. concentrated sulphuric acid. Pass the solution through the reductor and follow with 200 cc. of water, *taking care that no air passes through the reductor*. It is easily shown, qualitatively, that hydrogen peroxide is formed by passing air and water alternately through the reductor. After standing, even for a few minutes, the reductor should be rinsed with dilute sulphuric acid before use. Blank determinations should be made frequently and the amount of

[1] The method given by Dudley and Pease (This JOURNAL, 15, 521,) has given me practically identical results with the above methods, also piano wire, considering it as pure iron, the impurities in iron, some of them, requiring more permanganate than iron itself. The following results were obtained:

Iron wire (Fresenius' method)	0.003199
Ammonium oxalate	0.003209
Ferrous ammonium sulphate	0.003209
Dudley's steel	0.003179
" "	0.003205
" "	0.003208

W. A. NOYES.

permanganate required should be subtracted from that found in an analysis. The blanks should not take more than two-tenths to three-tenths cc. of permanganate. Ours are usually less than two-tenths cc. We do not think it advisable to use the same reductor for both iron and molybdenum solutions.

ANALYTICAL PROCESS.

Weigh two grams of steel[1] to an accuracy of two mgms., put it in a 500 cc. (16 oz.) flask, add fifty cc. of nitric acid (1.18), and warm, only gently, till solution is complete, (our experience leads us to think some phosphorus may escape if the gases in the flask become colorless before solution is complete); then boil for one minute, add ten cc. of the permanganate solution (one and one-fourth per cent.), boil till the pink color disappears, add ferrous sulphate carefully, with shaking, till the solution clears, cool to about 50°, add eight cc. ammonia (0.90), insert the stopper, and shake till any precipitate which forms redissolves; cool or warm till the solution is as many degrees above or below 60° as the molybdic solution is below or above 27°, add sixty cc. of molybdic solution, insert the rubber stopper, and shake by hand or with a machine for five minutes; allow to settle for five minutes, filter through a nine cm. filter of rather fine texture, wash with the acid ammonium sulphate solution, putting five or ten cc. at a time into the precipitation flask and pouring from that all around the upper edge of the filter. The filtrate and washings must be absolutely clear. Wash till five cc. of the washings give no brown color with a few drops of dilute hydrogen sulphide water. Pour over the precipitate twelve cc. of ammonia (0.96) diluted to twenty-five cc., allow the solution to run into the precipitation flask, wash with 100 cc. of water, add eighty cc. more of water and ten cc. of concentrated sulphuric acid, pass the solution through the reductor and follow it with 200 cc. of water, *taking care that no air passes through the reductor* during reduction or washing. Titrate the solution, *at once, in the reductor flask.*

With steels high in phosphorus, the ammoniacal solution of the yellow precipitate may be made up to a known volume and an aliquot part taken for reduction.

[1] Dr. Dudley suggested the use of two grams of steel.

For cast iron low in phosphorus, the method may be carried out exactly as given for steel. The method may also be used without modification other than the use of a larger amount of ammonia for solution of the yellow precipitate and the use of an aliquot part of the solution for reduction, for cast iron high in phosphorus. We prefer, however, in that case, to use one gram of the iron, dissolve, oxidize, clear, and neutralize as usual, using twelve cc. of ammonia. Then the solution is made up to 150 or 250 cc. and ten cc., or more, taken, enough water added to make fifty cc., the solution warmed to 60° and sixty cc. of molybdic solution added. The remainder of the process is as above for steel. The residues insoluble in nitric acid may require examination as usual.

The process as here described has given us the results stated in the table below. No determinations which have been made by this method have been omitted. Steels Nos. 1, 2, and 3 were kindly furnished us by Dr. Dudley, and are the same as those described in this JOURNAL, **16**, 217, and the gravimetric determinations given are the very careful ones reported in that paper. Steel "a" was also furnished us by Dr. Dudley with the statement that it contains, by their determination, 0.016 per cent. phosphorus, and is the lowest they have. Steel C is a tool steel containing, probably, about one per cent. of carbon, and, by our own gravimetric determination by the combined acetate-molybdate-magnesia method[1], using ten grams of the steel, 0.021 per cent. of phosphorus. Steel A is a mild steel containing about 0.17 per cent. carbon and, by the molybdate-magnesia method, using five grams, 0.129 per cent. phosphorus. We do not place as much reliance on this gravimetric determination as on the other.

The gravimetric determinations for the cast iron were made as follows: one gram of the iron was dissolved in fifty cc. of nitric acid (1.18) and treated exactly as in our directions for steel. The ammoniacal solution of the yellow precipitate was made up to 150 cc. and two portions of ten cc. each were taken for reduction. These portions gave determinations 18–21 of the table. The remainder of the ammoniacal solution was concentrated in a platinum dish and then precipitated with magnesia mixture.

[1] This JOURNAL, 16, 221.

The magnesium pyrophosphate, after weighing, was dissolved in hydrochloric acid and evaporated to separate silica. In the first determination no silica was found; in the second there was an appreciable amount. The gravimetric determinations gave 1.75 and 1.79 per cent. of phosphorus, respectively.

In order to determine if more phosphorus could be obtained by solution in aqua regia, one gram of the iron was dissolved in thirty cc. of a mixture of equal volumes of concentrated nitric and hydrochloric acids; the solution was made up to 150 cc.; ten cc. were placed in a porcelain dish with five cc. of concentrated nitric acid, evaporated to dryness, and heated to 200° for one-half hour, the residue taken up with fifty cc. of nitric acid (1.18), nineteen cc. of ammonia added, and the solution precipitated as usual. This method gave the determinations 22 and 23 of the table. Determinations 24 and 25 of the table were made by the method recommended above for cast iron.

		Gravimetric determination.	Volumetric determination.
1.	Steel No. 1	0.040 per ct.	0.037 per ct.
2.	" " 1	0.040 "	0.040 "
3.	" " 1	0.040 "	0.041 "
4.	" " 1[1]	0.040 "	0.0404 "
5.	" " 2	0.053 "	0.0512 "
6.	" " 2	0.053 "	0.0548 "
7.	" " 2	0.053 "	0.0526 "
8.	" " 3	0.032 "	0.0317 "
9.	" " 3	0.032 "	0.032 "
10.	" " "a"	0.016 "	0.0177 "
11.	" " "a"	0.016 "	0.0174 "
12.	" " C	0.021 "	0.021 "
13.	" " A	0.129 "	0.129 "
14.	" " A	0.129 "	0.129 "
15.	" " A	0.129 "	0.127[2] "
16.	2 gram steel No. 1 + 0.0173 gram. cryst. sodium phosphate	0.115 "	0.116 "
17.	2 gram steel No. 3 + 0.0231 gram cryst. sodium phosphate	0.132 "	0.133 "
18. 19.	1 gram cast iron	1.75 "	1.743 " 1.767 "
20. 21.	1 " " "	1.79 "	1.761 " 1.782 "
22. 23.	1 " " " dissolved in aqua regia	"	1.755 " 1.743 "
24. 25.	1 " " " " as for steel	"	1.750 " 1.745 "

[1] Determination by W. A. Noyes.
[2] With fresh, colorless, molybdic solution.

These determinations furnish, we believe, satisfactory proof that the method which has been described, secures a practically complete precipitation of the phosphorus, even when the amount present is only 0.02 per cent., and also that the ratio of phosphorus to molybdenum in the solution is constant and independent of the amount of phosphorus present; also that the ratio is not affected by the presence of silicon or arsenic in the solution. The cast iron used contains 0.84 per cent. of silicon. As regards arsenic, we have applied Marsh's test for arsenic to the yellow precipitate obtained from two grams of steel No. 2, and obtained only a very faint arsenic mirror, much fainter than that obtained under the same conditions, from 0.02 mgm. of arsenious oxide. If we suppose the arsenic to have been present in a compound analogous to the phosphorus compound, the molybdenum associated with it would correspond to only 0.0003 per cent. of phosphorus. The determinations with the cast iron also furnish evidence that the conversion of phosphorus into phosphoric acid is as complete by solution in nitric acid and oxidation with potassium permanganate as by solution in aqua regia and evaporation with nitric acid. The volumetric determinations with the cast iron are more accordant than the gravimetric. In view of the great difficulties known to exist in the way of securing pure magnesium pyrophosphate, we believe the volumetric determination will usually be found to be more accurate.

TERRE HAUTE, December 20, 1894.

THE CONTRIBUTIONS OF CHEMISTRY TO THE METHODS OF PREVENTING AND EXTINGUISHING CONFLAGRATION.[1]

BY THOMAS H. NORTON.

Received December 24, 1894.

I have felt it desirable on this occasion to direct your attention to a brief review of the services which the chemist has thus far rendered in the battle with fire, to the field of investigation still open, to the methods of testing and experimentation in this branch, and to the application of fact already gained to American conditions. In the choice of this subject, I have, in common

[1] Abridged from the author's opening address before the Section of Chemistry of the American Association for the Advancement of Science, at the Brooklyn Meeting, August, 1894.

with some of my predecessors, left the domain of pure science to consider more closely certain economic aspects of chemistry, and enter, to some extent, into the province of chemical engineering, which is now recognized in our leading schools as entitled to distinct and separate treatment. A further reason which has influenced me is the practical absence, to this time, in our works of reference, of any attempt to collate and classify the knowledge gained in this field.

The total annual loss of insured property by fire is about $200,000,000, and of this sum nearly one-half occurs in the United States,—$90,344,075 in 1893. Foreigners marvel at this reckless disregard of the immense losses by fire shown by our countrymen, for from eighty to ninety per cent. of the 15,000 fires which occur annually, can be traced to easily preventable causes, and chiefly to faulty construction. Consider, for a moment, the tax which is being paid this Moloch. Our total loss of property amounts to one-fifth of the net profits of all the industries of the country. Propose to a trader or manufacturer cooly to throw one-fifth of the annual profits into the grate! Yet this is what our productive industries, as a whole, have been and are doing in an unconcerned, if not cheerful, manner. The direct loss by conflagration is, however, not the only factor. Our fire departments and water supply cost us $30,000,000 annually; while the loss of wages to mechanics and other productive forces, and numerous contingent amounts, swell the actual total loss to a most serious sum. The fire department of London costs but one-third of that of New York City; and the same ratio applies to most European cities. During a residence of ten years in French and German cities, I saw the fire engines called out but five times; while the average resident of an American city is apt to witness a call at least once a week. England's annual fire bill is but little over $10,000,000.

With these facts in view, is it not time to call a halt, to bring all the forces of science to bear in the battle, and redeem our America from what is nothing more or less than a servile bondage?

If we seek the reason for this vast difference in relative fire loss between the Europe of to-day and America, it is not to

be found in a superiority of facilities for extinguishing conflagrations. The equipment of American fire departments is far better than that of the fire departments abroad. It results simply from the fact that the European architect and builder have profited from the experience of the past, and learned that prevention is better than cure. The prevalent method of construction in New York is such, architects assure me, that the business part of the city awaits only the proper combination of circumstances to be the scene of a conflagration rivalling that of Chicago.

Before leaving these general considerations, it should be said that there are many hopeful indications of an awakening from this apathy, and ample recognition should be made of the earnest effort, inaugurated of late years in New England, to lessen fire risk ; the application of the so-called slow-burning principle of construction, especially to factories. By the adoption of this principle it has been possible to reduce insurance rates from one and one-half to two and one-half per cent. down to two-sevenths of one per cent.

FIRE EXTINGUISHERS.

Passing to the more specific treatment of the subject, we find that, as in medicine, so in the matter of protection against loss by fire there are two distinct lines of action : prevention and cure, or the adoption of such measures as to render an outbreak of fire difficult, and a conflagration practically impossible ; and the provision of the proper facilities for the rapid subjugation and limitation of fire under full headway. As in medicine, also, there is a growing conviction that the prophylactic treatment is the more important. It is hence particularly in this direction that the activity of the chemist has been chiefly called into play.

Let us first briefly review the methods of extinguishing fire.

In this field but little has been done to add to the efficacy of the agent used from time immemorial—ordinary water. In addition to its cheapness and universal distribution, water possesses, over other liquids, peculiar advantages for the purpose in view—high specific heat and the formation of a vapor which is non-poisonous. Sulphur dioxide has been used to some extent; in fact it is an old-fashioned method in European countries to

extinguish burning chimneys by kindling sulphur on the hearth. A quicker production of the gas is effected, more particularly for use in confined spaces, by introducing receptacles of burning carbon disulphide; and cylinders of liquid sulphur dioxide, have, under similar circumstances, been successfully employed. The advantage in the use of this gas consists entirely in its exclusion of the air necessary to maintain combustion. The slower diffusion consequent upon high density renders it superior for this purpose to other available gases. This advantage is, however, largely counterbalanced by its irritant, poisonous nature, and it has never been accepted as one of the recognized agencies of general application.

Of much greater value is carbon dioxide, the efficacy of which is likewise based upon its ability to prevent the access of air to the material in process of combustion. The ease with which this is accomplished is shown by the familiar experiments of our lecture courses in chemistry. For practical results in dealing with ordinary fires, it is found best to employ a highly charged aqueous solution of the gas. Such solutions as those prepared for use as a beverage, have frequently been used in emergencies. One large soda-water establishment has been able to extinguish several fires in its own factory and in the neighborhood by the prompt use of the carbonated water in its receptacles. A quick-witted pharmacist recently extinguished a serious fire in his store from benzene, by quickly using a pailful of soda-water drawn from his own fountain, the case being one in which simple water would have failed to accomplish the purpose.

The so-called chemical fire-engines, now regularly used in our public fire departments, are all constructed upon the principle of charging a quantity of water when used with carbon dioxide and ejecting a stream of the carbonated water by the pressure of the gas itself. In the size most frequently employed the supporting truck carries two cylindrical tanks of steel or copper, holding about eighty gallons of water and twenty-eight pounds of sodium bicarbonate with a leaden jar containing fourteen pounds of sulphuric acid. At the moment of using, by a simple mechanical device, the acid is admitted to the solution and a pressure of 140 pounds to the square inch is developed. The

tanks are charged alternately while in operation so that it is possible to maintain a continuous jet of one-fourth of an inch. The whole apparatus is light and easily drawn by two horses. There is also in use a combined manual and chemical fire engine devised by Foster. The pumps fill, with water, the compartments in which the chemicals are dissolved and whence the solutions come together in a generating chamber where the gas is evolved. Such engines are capable of throwing per minute thirty gallons of water containing 250 gallons of carbon dioxide to a distance of ninety feet. This type is also arranged for connection with high pressure mains. Very careful tests carried out in 1883 showed conclusively the great value of a small quantity of carbonated water in extinguishing fierce flame, especially in a confined situation and when combustion had not penetrated much beneath the surface. The same principle was applied earlier by Baragwanath and others in a variety of forms, to portable extinguishers. Dick's extinguisher, largely used in England, is a cylindrical tank of seven gallons, easily carried on the back, giving a pressure of from 70 to 120 pounds, and possessing a projectile range of fifty feet. The peculiar feature of this type is the storing of the sulphuric acid in a glass container, which is easily broken when the apparatus is to be used. Zabel's apparatus, like Dick's, is charged with sodium bicarbonate, but the acid used is contained in a cylinder from which it is liberated by simple inversion. Devices, essentially the same, have been introduced by Masnata and VanWisker. In Carter's extinguisher the sulphuric acid is replaced by tartaric acid.

Platt's extinguisher has been successfully used for many years, and is valued on account of its simplicity; the turning of a valve and the inversion of the apparatus bringing it at once into operation. Other efficient extinguishers of American origin are the Babcock, the Harkness, and the "Climax." In the latter, sulphuric acid is replaced by oxalic acid, and in both the Harkness and "Climax" the carbonated solution is forced out by a pump. The Babcock extinguisher is filled with a solution of sodium bicarbonate, and has in its upper part a vessel of acid suspended by lateral pivots to a stirrup depending from the top of the apparatus. The stopper of this vessel is worked by a rod

through the top of the extinguisher. By withdrawing the stopper, the vessel tilts over and mingles the acid with the solution producing at once the necessary pressure.[1]

The most recent and perhaps most efficient of chemical extinguishers is that introduced in 1893 by Dr. Henry P. Weidig, and extensively tested throughout our country during the past twelve months. It is adapted not only for portable use, but also for replacing ordinary town engines and for permanent installation in factories and on board ships, etc.

The apparatus consists of a steel vessel containing liquid carbon dioxide so arranged that it may be brought in connection with a receptacle filled with carbonated water, or a solution of ammonium carbonate, under any desired pressure. The carbonated water meets in its exit pipe with a stream of ammonia gas emanating from a cylinder of liquid anhydrous ammonia. The carbon dioxide is thus bound in the form of a soluble salt, ammonium carbonate. The solution thus formed is forcibly ejected, and, on reaching a fire, the heat not only volatilizes the water but decomposes the salt, so that a mixture of steam, carbon dioxide, and ammonia is the result. One volume of water will yield under these circumstances twenty-five volumes of the two latter gases.

The principle of rapidly filling a confined space where fire has broken out with inert gases has been extended to the use of the gaseous products issuing from the combustion under boilers. In certain industrial establishments devices are introduced by which the current of the mixture of nitrogen, carbon dioxide, and aqueous vapor can be directed, at will, into a given apartment. The results are quite effective, combustion being promptly and permanently stopped. Other additions than carbon dioxide to the water used in extinguishing fires have proved of doubtful utility when submitted to practical tests. The list of such proposed additions includes potash (Kaiser), clay (Glaser), a mixture of sodium sulphite, and four parts sodium alum (patented 1884 in Austria, by Trotha, and sold in the form of cartridges), salt, alum, waterglass, copperas, Glauber's salt, borax, magnesium chloride, sodium phosphate, and soda.

[1] Hexamer, *Journal of the Franklin Institute*, August, *1885*.

Ammonia gas alone has been recommended. Its lightness and irritant properties render it, however, much less efficient than carbon dioxide or sulphur dioxide.

Hand grenades,[1] consisting of glass bottles, usually of a spherical form, charged with aqueous solutions of various chemicals, have found their way of late years into somewhat extended use. They are conveniently arranged in the corridors of large buildings, and are intended for use in extinguishing incipient outbreaks of fire. The grenade is to be thrown with such force into the center of combustion that it breaks and scatters its contents upon the burning material. The salts present are supposed to act by either giving off incombustible gases or by incrusting the burning articles, thereby excluding air. In practice, however, it is found that these grenades render but slight service, as they are frequently unbroken when thrown, especially against non-resistant substances. They are supposed to be charged with strongly carbonated water, or with a saturated solution of ammonium chloride, sodium chloride, sulphite, or thiosulphate. Analysis, however, shows quite a variety of composition in the contents. Sir Frederick Abel,[2] in 1881, found one to contain a strong solution of salt and waterglass. Geissler examined three current types with the following results:[3]

Hayward's Grenade.—Gross weight 1,120 grams; weight of liquid, 750 grams; composition, an aqueous solution of 15.7 per cent. calcium chloride and 5.6 per cent. magnesium chloride.

Harden's Grenade.—Gross weight, 900 grams; weight of liquid, 555 grams; composition, an aqueous solution of 19.5 per cent. sodium chloride, and 8.9 per cent. ammonium chloride.

Schonberg's Grenade.—Gross weight, 770 grams; weight of liquid, 440 grams; composition, an aqueous solution of 1.66 per cent. sodium carbonate, and 6.4 per cent. sodium chloride.

It may be noted in passing that the salts in these grenades cost about one per cent. of their selling price.

Finally, mention should be made of extinguishing powders,[4] the use of which is literally a fighting of fire with fire. They

[1] Hexamer, *loc. cit.*
[2] Amer. Arch, 13 and 14.
[3] Pharm. Centrbl., *1885*, 447.
[4] Hexamer, *loc. cit.*

are well devised to produce, by rapid combustion at the expense of the oxygen in saltpeter, a sufficient volume of inert gas—chiefly sulphur dioxide and carbon dioxide—to prevent the circulation of air about a fire, and thereby cut off the supply of atmospheric oxygen for further combustion. Their application is, of necessity, limited to small enclosed spaces without many openings for ventilation, and they have been found chiefly valuable in drying rooms where easily volatile products are being driven off, as in connection with the manufacture of oilcloth. The ingredients are coarsely powdered and readily ignited. They are often packed in cartridges and provided with fuses. They burn without explosion, with a bright white flame, a strong odor, and much smoke. Each pound yields about four and eight-tenths cubic feet of gas, which consists of two and three-tenths sulphur dioxide, one and one-tenth parts carbon dioxide, and one and four-tenths parts nitrogen. One pound should be used for every 240 cubic feet of an enclosed space. The composition of the best known of these powders is as follows:

Bucher's powder	KNO_3	60	parts.
	S	36	"
	C	4	"
Heeren's powder	KNO_3	63	parts.
	S	29	"
	C	4	"
	Fe_2O_3	4	"
Schweizer's powder	KNO_3	58	parts.
	S	36	"
	C	3	"
	Fe_2O_3	$1\frac{1}{4}$	"
	SiO_2 (sand)	75	"
Gruneberg's powder	KCl	20	parts.
	KNO_3	50	"
	S	50	"
	MnO_2	1	"
	Rosin	10	"
Zeisler's powder	KNO_3	60	parts.
	S	36	"
	C and CaO	4	"

Johnston's powder	Equal parts of KCl, KNO_3, MnO_2 and Rosin, moistened with waterglass and pressed into briquettes.
Phillips' Fire Annihilator (introduced in 1851.)	KNO_3 60 parts. C 20 " $CaSO_4 + 2H_2O$ 5 "

This latter was pressed in the form of a hollow brick. The central cavity contained sugar, and potassium chlorate and a vial of sulphuric acid, the latter so arranged as to be easily broken by a simple mechanical device. The brick was usually placed within a double cylindrical receiver containing water, so that the ignition of the powder caused not only the evolution of a large volume of gas, but also, by its heat, a liberation of a considerable amount of aqueous vapor.

In reviewing this portion of our subject it would appear that but few opportunities are afforded the chemist for further contributions of value. It would seem advisable, however, to study the feasibility of extending the use of carbon dioxide along lines similar to those in which automatic sprinklers have found such satisfactory application. In many factories and in numerous stores and warehouses the sprinkler, while most effective in quenching flame or hampering its progress, causes serious damage to manufactured products or materials for manufacture. The advantages of carbon dioxide over water under these conditions are manifest. It is not difficult to conceive of a system of piping, opening naturally at the level of the floor, through which either automatically or under the control of a watchman, a powerful current of the gas could be directed into any apartment where flame was apparent. In the many cases where the losses by water do and would naturally far exceed the losses by direct combustion, there is but little doubt that the use of the invisible, inert, harmless gas would enormously limit the destruction of valuable property.

We come now to the most important division of our subject; namely, the methods of preventing loss or suffering by fire through the use of structural materials, textile fabrics, and the like, which are either entirely incombustible or slowly combusti-

ble or from which the property of rapid combustion has been largely removed by special treatment. A structure built entirely of stone, brick, iron, cement, etc., is an illustration of the first type. A structure of brick with heavy beams of wood, heavy floors of plank, no hidden air-spaces, and a general absence of the lighter forms of woodwork, illustrates the second type,—the slow combustion construction. The third type is to be found in an edifice of wood, the component parts of which, either by impregnation or by suitable coatings, have lost the property of ready inflammability.

The study of the choice of materials and their most efficient grouping as employed in the first and second methods of construction, apart from architectural considerations, falls purely within the domain of the mechanical engineer. It involves the question of strains and stresses under ordinary conditions and under the conditions of an elevated temperature.

It is in the special field of rendering such easily combustible substances as wood and the vegetable fibers used in textile fabrics more or less resistant to flame that the chemist has been able to render service.

Could our methods of building be limited to the completely fire-proof or the slow-burning system of construction, and interior decoration and equipment be restricted to the use of non-combustible materials, it is evident that the chemist would have but little to do. But mechanical considerations, canons of taste and questions of outlay will, for a long time to come, sustain the extended use of wood in the erection and finishing of many classes of edifices, especially of our homes. The beginnings of our towns and cities are almost entirely in wood, and the same may be said of the suburbs of our centers of population. It will be many years before the rural residence and the home of the laborer in America will be as universally constructed of non-combustible materials as in Great Britain, France, Germany, and other European countries. In such a country as Japan it is almost impossible to imagine any general departure from the time-honored and picturesque national construction of bamboo. The restriction of fabrics to fibers of animal origin is likewise out of the question.

In noting the historical development of the work of the chemist in the field as above outlined, we will first consider the methods which have been applied to textile fabrics and paper, and next those which have been used to render wood uninflammable.

TREATMENT OF TEXTILES AND PAPER.

The first experiments in this direction were made in England in 1735 by Obadiah Wild,[1] who secured a patent for rendering the cartridge cases used by the navy resistant to flame. This process consisted of the addition of a mixture of alum, borax, and copperas, to the paper pulp employed. Early in the present century the subject attracted the attention of several chemists. De Hemptine, of Belgium, in 1821, made quite an extended study of mixtures, similar to that of Wild, and other substances. At the same time Brugnatelli recommended the use of potassium silicate, Hermbstadt of ferrous sulphate, De Lisle of an unnamed compound.[1] Gay Lussac[2] was deeply interested in the subject. In 1830, as the results of his experiments, he recommended the use of various chlorides, phosphates, borates, tartrates, and carbonates without, however, considering the difficulties involved in their application on a large scale. The employment of the alkaline carbonates was later urged strongly by Prater, while Fuchs[1] and Bethel in 1838 advocated waterglass and others, gypsum. Several patents were taken out during this period by the different investigators. Still later Chevalier recommended the use of a mixture of borax and ammonium sulphate. The use of ammonium sulphate alone was proposed by De Breza in 1838, and that of ammonium chloride by Froggant in 1857. In 1855 M. Solomon,[3] of Paris, submitted to the English Board of Ordnance a process for rendering canvas uninflammable. It consisted of immersing the materials for a day in each of the following baths:

I. Aluminum sulphate,	30 parts.	II. Dry calcium chloride,	20 parts
Glue (gelatin),	10 "	Glue,	10 "
Water,	60 "	Water,	70 "

1 Versmann and Oppenheim, Report to the British Association, *1859*
2 *Ann. chim. phys.*, 18, 211.
3 Amer. Arch., 13 and 14.

[TO BE CONTINUED.]

THE ACTION OF ORGANIC AND MINERAL ACIDS UPON SOILS.

By Harry Snyder.

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IN soil analysis, the hydrochloric acid use for dissolving the potash, lime, and phosphates has a specific gravity of 1.115. This acid is used because it is supposed to represent, practically, the same solvent power which agricultural plants possess in obtaining their food from the soil. Inasmuch as this premises is the foundation of a part of the present system of soil analysis, it is a question of the greatest importance. One of the indications which points to the truth of the premises is the similarity of the action of hydrochloric acid of this strength, and of organic acids upon samples of soil.

Organic Acids.—In some of the soil work which has been in progress in the chemical laboratory of the Minnesota Experiment Station, the solvent action of the various organic acids has been determined with another ultimate object in view. Incidentally, the results give sufficient data for the comparative action of the organic acids and of hydrochloric acid, sp. gr., 1.115, upon the same samples of soil.

Different types of soil are acted upon quite differently by the various organic acids, and general conclusions regarding their action upon various types of soil can not be stated. Hence the statements which are made upon this question have reference only to the type of soil under consideration.

The figures given in the table were obtained from a soil sample taken from one of the rotation plots at the Experiment Station. In obtaining the solutions for analysis, oxalic, tartaric, and citric acids, of ten per cent. strength, by weight, were used. The digestion was carried on for thirty-six hours in the Snyder soil flask as used at this station. The organic acid mixture was made by preparing a ten per cent. solution, by weight of citric, oxalic, and tartaric acids. The material soluble in the organic acids was separated by filtration and washing. The solution was then gently ignited to remove the organic acids, so that the analysis could be completed in the usual way.

Ten per cent. solutions were used because the more concentrated ones crystallized upon standing, and, then too, preliminary trials indicated that the more concentrated organic acid solutions possessed less solvent power than those here used. One difficulty experienced in the work was obtaining tartaric acid free from potash, and citric and oxalic acids free from lime. All of the determinations were made in duplicate, and controlled by blank determinations.

TABLE I.

SOLVENT ACTION OF HYDROCHLORIC AND ORGANIC ACIDS UPON A SOIL.

	Hydrochloric acid 1.115 sp. gr.	Ten per cent. solution of			
		citric acid.	oxalic acid.	tartaric acid.	organic acid mixture.
Total insoluble, per cent	84.08	87.90	85.07	87.67	85.93
Potash	0.30	0.12	0.18	0.06	0.10
Soda	0.25	0.17	0.19	0.05	0.12
Lime	0.51	0.50	0.23	0.41	0.45
Magnesia	0.26	0.27	0.21	0.32	0.33
Iron oxide	2.56	1.59	1.14	1.62	2.23
Alumina	4.24	1.45	3.96	1.77	2.99
Phosphoric anhydride	0.23	0.26	0.14	0.11	0.26
Sulphuric "	0.08	0.10	0.05	0.06	0.05

The oxalic acid caused an increase in the solubility of the alumina. There is a marked difference between the solvent action of these three acids. The oxalic acid possesses the greatest solvent action upon the complex insoluble matter of an alkaline nature. The citric acid possesses the greatest solvent action upon the lime, magnesia, and the sulphuric and phosphoric anhydrides. The amount of lime dissolved by the oxalic acid is greater than would, at first, be expected, but the complex nature of the solution, and the presence of the salts of other metals, has caused nearly half as much lime to be dissolved by the oxalic acid as by the hydrochloric acid of sp. gr. 1.115.

The ten per cent. mixture of the three organic acids has resulted in dissolving as much phosphates and magnesia, and nearly as much lime as any one of the acids acting alone. The potash does not appear to be as soluble in the organic acid mixture as the other elements of plant food.

The amount of potassium, iron, and aluminum oxides, dissolved by these three organic acids and their mixtures is somewhat less than that dissolved by the hydrochloric acid. The

iron and alumina are not of any economic importance in interpreting the results of soil analysis. The main difference between the solvent action of the hydrochloric and the organic acids is in the somewhat smaller amount of potash dissolved by the organic acids.

With the exception of the solubility of the potash, as noted, these results bring out, in a general manner, the value of hydrochloric acid of 1.115 sp. gr. as a solvent, over the organic acids, of ten per cent. strength, because the hydrochloric acid represents the limit of the solvent action reached by any one of the organic acids acting alone or in a mixture.

For the purpose of comparison, the solvent action of hydrochloric acid, sp. gr. 1.115, and the maximum limit reached in any case, by the organic acids, are given, together with the name of the acid possessing the highest solvent power for that constituent.

TABLE II.

	Hydrochloric acid, sp. gr. 1.115.	Largest amount dissolved by organic acids.	Name of the organic acid.
Total insoluble, per cent	84.08	85.07	Oxalic
Potash	0.30	0.18	Oxalic
Soda	0.25	0.19	Oxalic
Lime	0.51	0.50	Citric
Magnesia	0.26	0.27	Citric
Iron oxide	2.56	2.23	Mixture
Alumina	4.24	3.96	Oxalic
Phosphoric anhydride	0.23	0.26	Citric and Mixture
Sulphuric "	0.08	0.10	Citric

Mineral Acids.—The action of the concentrated mineral acids upon the same sample of soil is given in Table III. Concentrated hydrochloric acid, forty per cent. strength, dissolved two and a quarter per cent. more of the complex silicates than nitric acid of the same strength. Concentrated nitric acid possesses a stronger solvent action than the dilute hydrochloric acid. The concentrated hydrochloric acid dissolved somewhat more potash than the hydrochloric acid of 1.115 sp. gr. The amount of lime, phosphates, and sulphates, dissolved by the concentrated hydrochloric acid, was about the same as when the more dilute acid was used; the concentrated acid, however, dissolved more of the complex silicates of potash, iron, and alumina. Sulphuric acid possesses the strongest solvent action.

Of the total potash in the soil, as determined by fusion, less than eighteen per cent. is soluble in the hydrochloric acid as used in soil analysis; about ten per cent. is soluble in oxalic acid, and three per cent. in tartaric acid.

TABLE III.

	Concentrated Acids.			
	Hydrochloric acid.	Nitric acid.	Sulphuric acid.	Total by fusion methods.
Total insoluble, per cent	81.20	83.45	80.45	72.45
Potash	0.42	0.30	0.52	1.75
Soda	0.31	0.28	0.40	0.50
Lime	0.55	0.30	0.53	0.86
Magnesia	0.40	0.32	0.52	0.72
Iron oxide	3.41	2.97	3.57	3.63
Alumina	5.84	4.55	6.88	12.71
Phosphoric anhydride	0.23	0.23	0.26	0.28
Sulphuric "	0.08	0.08	0.10	0.10

NEW BOOKS.

A TREATISE ON INDUSTRIAL PHOTOMETRY, WITH SPECIAL APPLICATION TO ELECTRIC LIGHTING. BY A. PALAZ, PROFESSOR OF INDUSTRIAL ELECTRICITY IN THE SCIENCE FACULTY OF THE UNIVERSITY OF LAUSANNE. TRANSLATION BY GEO. W. PATTERSON, JR., ASSISTANT PROFESSOR OF PHYSICS IN THE UNIVERSITY OF MICHIGAN, AND M. R. PATTERSON. pp. 322. New York: D. Van Nostrand Co., 1894. Price, $4.00.

Many of the methods of photometry which have been long in use for measuring the intensity of light from gas flames have required modification when applied to the measurement of the intense light produced by electricity. A book which gathers together and arranges, systematically, the work on the subject which is scattered in many different journals, and which is frequently difficult of access, will be very welcome to many who are interested in the subject.

The first part of the book gives a discussion of the general principles of photometry and includes a description of all of the common photometers and photometric standards. This part of the book is quite full and will be of value to anyone interested in photometry in any of its applications. Then follow chapters on the equipment of photometric laboratories, and on the measurement of the intensity and distribution of light from electrical sources.

In an appendix of eight pages the translators criticise or give additional information upon a few topics in the book.

The work of translation appears to have been carefully done, and the English used by the translators is good. A short index is given, which might be made more comprehensive with great advantage. There are a great many short references to particular topics, and to the work of individuals, which can not be found by means of the index. W. A. NOYES.

DESCRIPTIVE INORGANIC GENERAL CHEMISTRY. A TEXT-BOOK FOR COLLEGES. BY PAUL C. FREER, M.D., PH. D. (MUNICH), PROFESSOR OF GENERAL CHEMISTRY AND DIRECTOR OF THE LABORATORY OF GENERAL CHEMISTRY, UNIVERSITY OF MICHIGAN. Revised edition. pp. 550. 54 Illustrations. Boston: Allyn and Bacon. 1894. Introductory price $3.00.

This book is intended as a text-book for colleges and is therefore adapted for the use of students who already have some knowledge of the elementary principles of the science. The treatment of the subject is consequently somewhat different from that usually pursued in the best of the recent text-books on chemistry. For example the book begins with a discussion of the atomic theory and the composition of chemical compounds. The author believes that "beginners should be taken through a course in which only a few elements and compounds are discussed, with the purpose of familiarizing the pupils with the fundamental laws which govern chemical change." "During the progress of such work as this," he continues, "I would not advise the use of chemical symbols or any reference to the atomic theory. It is manifestly impossible to make a student without experimental knowledge, understand, in all its bearings, a theory which it has taken some ninety years to place upon its present footing. If an elementary course, in which every stated fact has been proved by actual experiment, precedes the work given in this book, the pupil will be then amply fitted to look at chemical phenomena from the basis of the atomic theory."

Professor Freer is undoubtedly right in this matter and the sooner the teachers of chemistry in this country recognize the truth contained in the above-quoted statements and model their instruction on the plan there suggested the better it will be for the teaching of the science of chemistry in our colleges and uni-

versities. If the teaching of the elements of the subject could be relegated to the secondary schools, where it properly belongs, then it would be possible to give a very much more satisfactory course in chemistry in our colleges and universities on the lines laid down by Dr. Freer in the present book.

In the descriptive part of the work the author has used the last editions of Graham-Otto's "Lehrbuch der Allgemeinen Chemie," and Ladenburg's "Hardwörterbuch der Chemie," though the original articles in the chemical journals have been freely consulted. The elements have been discussed, in general, in the order in which they occur in the periodical system and the attempt has been made to show the influence exerted by the nature of the elements which form a chemical compound.

The book must be regarded from all points of view as a very good one and one that can be highly recommended for use in our colleges and universities. The facts are stated clearly, concisely, and correctly, and the latest information obtainable on the subject is always presented. Teachers of chemistry will find the book full of new ideas and suggestive of the relations of the various phenomena to one another that are usually overlooked in the ordinary text-book on the subject.

What the author has to say on the subject of valence in the preface to the book is very interesting. He says: "My views upon the subject of valence and the use of structural formulas may probably be regarded as too conservative by many of my colleagues, but I have been led to adopt these views by the growing conviction that the dogmatic use of supposed laws of valence and of constitutional formulas founded upon very incomplete experimental evidence, is causing more harm than good to the advancement of chemical science." Most of the teachers of chemistry will endorse the latter part of this statement. The author has made free use of the facts derived from the domain of the so-called physical chemistry wherever he has thought it advisable, and he has thus been able to present reasons for our views of certain chemical phenomena which would be otherwise perfectly inexplicable to the student.

Notwithstanding the general excellence of the book it has some minor faults. The use of the word not-metal throughout

the work when the word non-metal is in general use by chemists seems ill-advised and unwarranted. The same statement may be made also with reference to the word not-conductor. Hydroiodic acid and stibional are used instead of the more usual hydriodic acid and antimonyl and for no apparent reason. Hydrazine, hydroxylamine, azoimide, and methine are written without the final *e* probably because that is the German way of spelling these words and the author is more familiar with them in German than in English. No reference whatever is made to the preparation of the diamond artificially by Moissan and, in speaking of the preparation of the acid HN_3, the simplest and best method of all that have been proposed, *viz.*, that of Wislecenus, by the action of nitrous oxide on sodium amide, is entirely omitted.

On page 278, near the bottom, occurs the following remarkable statement: "By a judicious combination of the iodides of organic radicles, carbon and hydrogen compounds containing as many as sixty *carbonates* in a molecule have been prepared." In a foot-note, Hill and Hägele are made responsible for this "judicious combination," but it is doubtful if these chemists would recognize their own statement in this form.

HN_3 is referred to as azoimide or hydrogen nitride and the salts of the acid are called nitrides. It would seem to the reviewer that a name which recalled the fact that the substance was an acid, would be preferable, and, since there are quite a large number of nitrides known which are certainly not salts of the acid HN_3, much unnecessary confusion will be introduced by the use of the word nitride for a salt of the acid HN_3. These are, however, but slight faults which may be readily corrected in the next edition of the book and that another edition of so excellent a work will soon be called for seems highly desirable and also highly probable.

The book is singularly free from typographical errors, is well printed and well bound, contains a useful appendix of laboratory notes intended as a guide to teachers and students in compiling a list of experiments, and last, but by no means least, a very complete index.

W. R. ORNDORFF.